A

C ass-Book of Science

PART II.

CHEMISTRY

BY

PREM SINGH, M.Sc.,

Principal, Government Intermediate College, Campbellpore

AND

RAM BHEJA MAL SETH, M.Sc.,

Lecturer in Chemistry, Government College, Lyallpur.

REVISED BY

SAIN DAS, M.A., B.Sc. (*Cantab.*),

Formerly Principal, D. A. V. College, Lahore.

LAHORE

UTTAR CHAND KAPUR & SONS

1937

Printed by
L. Guranditta Kapur,
at
The Kapur Art Printing Works,
Abbott Road, Lahore,
and
Published by
R. S. Ram Jawaya Kapur,
Proprietor,
Uttar Chand Kapur & Sons,
Lahore.

PREFACE.

This little book forms Part II of ' A Class-Book of Science ' intended for the use of students preparing for physical science for the Matriculation and School Leaving Certificate Examination of the University of the Punjab. In the interest of students, we have considered it advisable to present Physics and Chemistry in two separate parts.

Our chief aim in writing the book has been to discuss the subject in a simple, logical and systematic manner so that the student not only understands easily what is placed before him but also follows it with interest. We have for purposes of illustration given, wherever possible, the names of places where certain substances are obtainable in our own country. We have also given their Indian equivalents and indigenous methods of preparation in the hope that this will arouse a genuine interest in the study of physical science, and would at the same time remove a certain amount of lamentable ignorance prevailing amongst our students about the resources of our own country.

Experiments form a substantial part of the text. A teacher of an experimental science succeeds in imparting knowledge far more convincingly by the help of experiments than he could by mere theoretical lessons. Nothing impresses the young mind more than the demonstration of an experiment carried on right before his eyes. It is therefore expected that in spite of the limited time and insufficient material at the disposal of an average school teacher, most of the experiments

which are simple and require very little of special apparatus, will be showr in the class-room. If this hope is realised, the authors feel convinced that the interest of the student in the study of the subject will become more real and useful than it is at present.

Our thanks are due to Messrs. P. Carter Speers, B. Sc., and N. A. Yajnik, M. A., A.I.C., Professors of Chemistry, Forman Christian College, Lahore, for their kindly going through the manuscript and making valuable suggestions.

PREFACE TO THE NINTH EDITION.

The book has been thoroughly revised and several additions and a few alterations have been made in the subject-matter.

Questions in Non-metals and Metals set in the M. & S. L. C. Examination of the Punjab University during the last ten years have been given at the end of Non-metals and Metals to give the students an idea of the type of questions that are usually set in the examination.

In order to arouse the interest of science teachers as well as students for some useful hobbies a new chapter on 'Scientific Hobbies' has been added in the form of Appendix at the end of the book.

We hope that these improvements will make the book more useful.

PREM SINGH
R. B. SETH

January, 1937.

CONTENTS

CHEMISTRY.

Chapter		Page
I.	INTRODUCTION	1
II.	SIMPLE CHEMICAL OPERATIONS	7
III.	ELEMENTS AND COMPOUNDS	19
IV.	AIR	29
V.	NITROGEN AND OXYGEN	36
	OXYGEN	38
VI.	WATER	46
VII.	HYDROGEN AND COMPOSITION OF WATER	56
	COMPOSITION OF WATER	62
VIII.	ACIDS, BASES AND SALTS	68
IX.	CARBON	74
X.	OXIDES OF CARBON	83
	CARBON MONOXIDE	91
XI.	COMBUSTION AND FLAME	95
XII.	HYDROCHLORIC ACID AND CHLORINE	102
	CHLORINE	106
XIII.	SULPHUR	113
XIV.	COMPOUNDS OF SULPHUR	121
	SULPHURIC ACID	124
XV.	COMPOUNDS OF NITROGEN	131
	AMMONIA	131
	NITRIC ACID	134
XVI.	PHOSPHORUS AND ITS COMPOUNDS	143
	COMPOUNDS OF PHOSPHORUS	148
	PUNJAB UNIVERSITY QUESTION PAPERS, NON-METALS	151

METALS.

		Page
XVII.	GENERAL PROPERTIES	157
XVIII.	SODIUM AND POTASSIUM	162
	POTASSIUM	169
XIX.	COPPER AND SILVER	174
	SILVER	177
XX.	ZINC AND MERCURY	180
	MERCURY	181
XXI.	ALUMINIUM	184
XXII.	TIN AND LEAD	187
	LEAD	189
XXIII.	IRON	193
XXIV.	ACTION OF ACIDS ON METALS	197
XXV.	SYMBOLS, FORMULÆ AND EQUATIONS	201
	PUNJAB UNIVERSITY QUESTION PAPERS, METALS	208
	APPENDIX, SCIENTIFIC HOBBIES	i

A

Class-Book of Science

PART II.

CHEMISTRY.

A
Class-Book of Science

CHEMISTRY.

CHAPTER I.

INTRODUCTION.

1. **Object of Chemistry.**—Matter exists not only in different forms, but it is also undergoing a continuous change; and it is the object of chemistry to study these changes and also the mutual action of different forms of matter in the universe. The knowledge of chemistry is not based on mere observations but is the result of experiments performed under various conditions. In short, **Chemistry** is that branch of science which enables us to understand the nature, composition and properties of different kinds of matter.

Physics and chemistry are so closely related to each other that it is difficult to draw a line of demarcation between the two. Generally speaking, **Physics** deals with the universal properties of matter, such as density, hardness, etc., and merely physical changes that are brought about by heat, electricity, etc. and that are more or less temporary in nature. **Chemistry,** however, treats of the composition of matter and of the phenomena which occur when different forms of matter react with one another.

2. Physical and chemical changes.—The changes that matter can undergo may cause a temporary or a permanent alteration in its properties.

Experiment 1.—Take a rod of iron with a wooden handle (an iron file with a handle will do) and place the end in the flame of a spirit lamp or that of a Bunsen burner. The end gradually becomes hot and then red hot. Take the rod out of the flame and allow it to cool. The above changes are reversed, and finally the rod returns to its original condition.

Experiment 2.—Heat a little sulphur in a test-tube; it melts to a yellow liquid. Continue heating; it darkens in colour and becomes too thick to be poured out of the tube. On further heating, it again changes to a thin liquid of a dark-brown colour. Stop heating and see that the above changes are reversed as the liquid cools.

NOTE.—*Do not heat sulphur so strongly as to drive out its vapour from the tube.*

Experiment 3.—Take a bar of soft iron, place it in a piece of glass tubing, and wind a silk-covered copper wire round the tube (Fig. 1). Pass an electric current through the wire. The iron bar will be found to have acquired the property of attracting small pieces of iron. Stop the current; the bar loses its

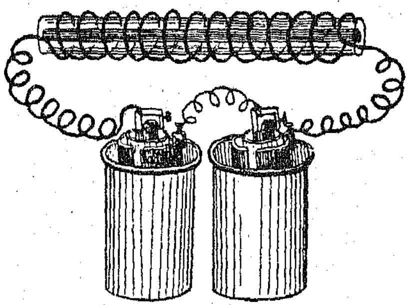


Fig. 1.—Magnetising iron by electric current.

magnetic properties.

In all the above cases the material of the substance remains unaltered, and the substance easily returns to the original state. Such changes that do not involve any alteration in the composition of matter are known as **physical changes.** The melting of ice into water and the conversion of water into steam are other familiar examples of physical change.

Experiment 4.—Hold a piece of wood or paper in a flame. It soon begins to burn and finally disappears, leaving behind a small residue.

Experiment 5.—Place a piece of magnesium ribbon in a flame. It burns with a dazzling light and leaves behind a white residue.

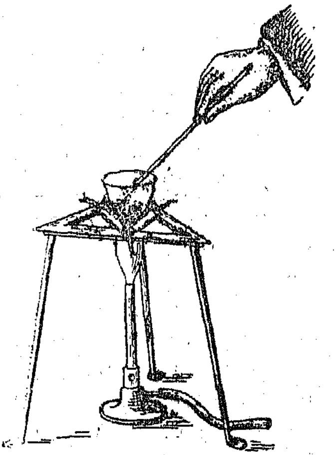


Fig. 2.—Heating of lead in air.

Experiment 6.—Heat some pieces of lead in a porcelain crucible placed on a pipe-clay triangle (Fig. 2); the lead melts. Stir the molten lead with an iron rod. The metallic lustre of lead gradually disappears and an orange-coloured substance (litharge) is formed.

In the above cases, the residue left behind cannot be easily changed back to the original substance and is entirely different in composition and properties from it.

When iron is exposed to moist air, it gets coated with a reddish-brown substance known as **rust,** which has entirely different properties from

those of iron. Similarly, when a candle burns in air, the products formed (carbon dioxide and water) are quite different from the candle. The rusting of iron and the burning of a candle are therefore chemical changes. Such changes that involve a distinct and a more or less permanent alteration in the form, composition and properties of matter are termed **chemical changes.**

The points of difference between physical and chemical changes may be summarised as given below:—

No.	Physiacl Change.	Chem cal Change.
1.	Merely a change of state from solid to liquid or liquid to vapour, and *vice versa*.	Not merely a change of state.
2.	No new thing formed.	New thing or things formed.
3.	No change in weight.	No change in the total weight of substances, but weights of individual substances generally change.
4.	Not accompanied by heat or light.	Generally accompanied by heat and sometimes light.

3. **Indestructibility of matter.**—Although a physical or chemical change can produce a temporary or a permanent alteration in the form and properties of matter, yet it can neither create nor destroy matter. From superficial observations one may conclude that the growing of a tree, the evaporation of water, and the burning of a candle prove the creation and destruction of matter. A careful study of these and numerous other phenomena by actual experiments has, however, shown that this gain or loss of matter is only apparent and not real. The tree grows because it receives nutrition from air and the soil; the water and the candle disappear because the vapours of water and the products of the burning candle pass into air.

Experiment 7.—Take an evaporating dish and weigh it. Put some water in it and weigh again. Now add a weighed quantity of sugar to the water and weigh the whole again. The weight of the dish and its contents will be found to be equal to that of the dish, water and sugar. Heat the dish by steam arising from water boiling in a beaker (Fig. 3) till all the water is driven out. Weigh the dish along with the dry residue. The residue will be found to be sugar, and equal in weight to the sugar originally taken. The sugar has undergone a temporary change, but it has suffered no gain or loss in weight.

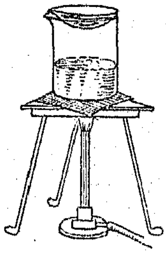


Fig. 3.—Heating a dish by steam.

Experiment 8.—Take a small flask fitted with a cork and a thin iron rod (Fig. 4). Put some dry sand in it and place a dry piece of phosphorus on the sand. Weigh the flask with its contents. Now remove the cork and heat the lower end of the iron rod in a flame. Quickly replace the cork and gently push the rod down till its hot end touches the phosphorus. The phosphorus catches fire and the flask is filled with dense white fumes. Allow the flask to cool, and weigh again. The weight remains unaltered. A chemical change has taken place, but matter has neither been created nor destroyed.

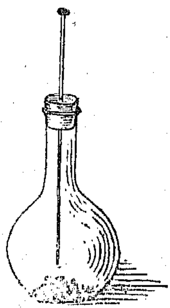


Fig. 4.—Burning of phosphorus.

Many experiments have been carried out, all proving directly or indirectly that matter can neither be created nor destroyed. This is known as the **Law of conservation or indestructibility of matter**, and may be stated thus:—*Matter can neither*

be created nor destroyed by any known physical or chemical means, although it may pass through many different forms.

QUESTIONS.

1. What is meant by a physical and a chemical change? State clearly the points of difference between the two. Illustrate your answer by at least two examples.

2. Classify the following as physical or chemical changes:—
(1) cooking of an egg, (2) burning of a candle, (3) heating of platinum wire, (4) burning of magnesium ribbon, (5) melting of ice, and (6) rusting of iron.

3. State, giving reasons, whether the changes involved in the following cases are physical or chemical:—
(1) Action of heat on sugar, iron and zinc.
(2) Burning of sulphur, kerosene oil and methylated spirit.
(3) Cooking of bread and vegetables.
(4) Melting of ice, butter and wax.
(5) Freezing of water and ice-cream.

4. What is meant by the indestructibilty of matter? Describe an experiment to show that matter is indestructible.

5. What chemical and physical changes are observed to take place when the following substances are strongly heated in open air:—(1) a crystal of common salt, (2) a crystal of alum, (3) a bright piece of copper, (4) a piece of magnesium wire, (5) a piece of sodium, (6) a piece of chalk, (7) powdered sulphur, and (8) ammonium chloride? [P. U. Matric., 1917.]

6. Distinguish as clearly as you can between a physical and a chemical change. Which of the following do you think are chemical and which are physical:—
(*a*) rusting of iron, (*b*) freezing ice-cream, (*c*) magnetising iron, (*d*) burning a candle, (*e*) burning an electric lamp, and (*f*) distilling water. Give reasons for your answers. [P. U. Matric., 1924.]

CHAPTER II.

SIMPLE CHEMICAL OPERATIONS.

4. **Solution.**—When sugar is added to milk and stirred, it disappears and imparts its sweet taste to the milk. Similarly, when common salt or alum is stirred with water, it spreads through the entire mass of the water and gives its own taste to it. Common salt or alum is said to be *dissolved* in water, and the process of dissolving as also the product obtained is termed *solution.* The substance which dissolves is called the *solute,* and the liquid in which it dissolves is known as the *solvent.* A **solution** is generally a clear transparent liquid in which the solute is so thoroughly mixed that it cannot be distinguished from the solvent. Every part of a solution has the same properties and composition as every other part, *i.e.* a solution is homogeneous.

Common salt, as stated above, readily dissolves when added to water. If, however, the addition be continued at a constant temperature a limit is reached when no more of it dissolves and if any more is added, the excess will remain in the solid state at the bottom of the vessel. Such a solution that can hold no more of the solid at a given temperature is said to be a **saturated solution.** On the other hand, a solution which contains less of the dissolved substance than it can hold in solution at that temperature is called an **unsaturated solution.**

5. **Evaporation.**—When a liquid is heated in a vessel, say a porcelain dish placed on a wire gauze held on a tripod (Fig. 5) or it is exposed to air in an open vessel, it gradually disappears. This process by which a liquid disappears on exposure or on heating is known as **evaporation.** It is generally used in recovering solids dissolved in water or other solvents. The solvent evaporates and the dry residue is left behind.

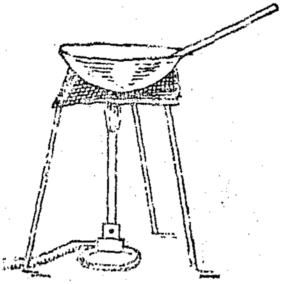


Fig. 5.—Evaporation by heating.

6. **Solubility.**—Some substances, such as sugar, nitre, common salt, etc., dissolve readily in water; others, such as lime, chalk, etc., dissolve only slightly; and still others, such as sand, glass, copper, etc., apparently do not dissolve at all. The amount of a solid that can be dissolved by water varies not only with the nature of the substance but also with the temperature. Most of the solids are more soluble in hot than in cold water. Water can dissolve about five times as much of nitre at 75°C. as at 20°C.

Experiment 9.—Take some water in a beaker and add nitre, a little at a time, with constant stirring. Continue adding till no more dissolves and some of it is left at the bottom. Now heat the solution; the solid at the bottom disappears, and a good deal more must be added before the solution is saturated. Allow the hot saturated solution to cool; a part of the solid separates out.

In the case of certain substances, such as lime and sodium sulphate, the amount of the solid dissolved decreases with the rise of temperature.

Experiment 10.—Take some burnt lime and prepare its cold saturated solution. Filter and heat the clear filtrate. It turns milky, showing the separation of the solid on heating.

The amount of a solid that can be dissolved in 100 parts of water at a certain temperature is known as the **solubility** of the solid at that temperature, and is determined as follows:—

Experiment 11.—Take some water in a beaker and add a little of powdered common salt to it. Stir well; if it dissolves, add more solid and stir. Go on adding till no more dissolves even after stirring for some time and some solid is left at the bottom. Filter the solution in a clean, well-dried and weighed evaporating dish. Weigh the dish with the solution and heat it gently on a sand-bath, *i.e.*, a small iron tray containing sand. As soon as the liquid begins to spurt, heat it on a water-bath (Fig. 6) or simply by steam from water boiling in a beaker (Fig. 3). When the whole of the water has evaporated and the dry solid is left in the dish, let it cool and weigh again.

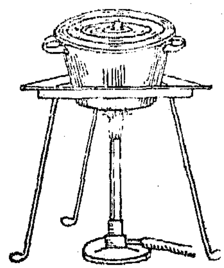


Fig. 6.—The water-bath

Suppose—

Weight of evaporating dish	=	46·49 gms.
Weight of evaporating dish and solution	=	70·31 gms.
∴ Weight of the solution.. ..	=	23·82 gms.
Weight of dish and residue	=	52·77 gms.
Weight of residue	=	6·28 gms.

And weight of water (= weight of solution *minus* weight of residue) = 23·82 — 6·28 gms.
= 17·54 gms.

i.,e. 17·54 gms. of water dissolve 6·28 gms. of salt at the room temperature.

Hence 100 gms. of water would dissolve
$$\frac{6{\cdot}28 \times 100}{17{\cdot}54} = 35{\cdot}8 \text{ gms. of the salt.}$$

Or, the solubility of common salt at the room temperature is 35·8.

Gases are, as a rule, less soluble in hot than in cold water. Air, carbon dioxide and other gases dissolved in water are driven out when the water is boiled.

7. **Water and other solvents.**—Water is said to be a universal solvent. It can dissolve something, however small, of every thing. Being the cheapest and the most common liquid, it is most frequently employed as a solvent. But there are other liquids which are also used as solvents, *e.g.*, alcohol, ether, benzene, carbon disulphide and petrol. A substance may be very slightly soluble in water but freely soluble in some other solvent. It is in these cases that other solvents are employed. Thus, sulphur and phosphorus are almost insoluble in water but soluble in carbon disulphide. Oils and fats are insoluble in water but dissolve in alcohol, benzene, or petrol. Hence it is that these liquids are often used to remove oil and grease spots from silk and woollen garments.

8. **Decantation.**—If a substance like sand or powdered marble is stirred up with water and allowed to stand, it settles down quickly. The clear liquid may be poured out or siphoned off without disturbing the solid. This process is known as **decantation.**

Experiment 12.—Stir a mixture of common salt and sand with water in a beaker, and allow the liquid to stand. The salt dissolves in water and forms a clear solution, while sand settles down quickly. Gently pour off the

clear liquid. Put some more water into the beaker and stir. Pour off the clear liquid again. By repeating the process, practically the whole of the salt can be removed from the sand.

In villages, drinking water containing sand and other insoluble impurities is allowed to stand and the clear liquid decanted off. The process is also used by *niyariyas* for separating gold and silver particles from the lighter impurities.

9. **Filtration.**—Decantation is not a satisfactory process for the complete separation of a liquid from the insoluble or suspended impurities, hence filtration is the method usually employed. The process of filtration consists in allowing the solution of the solid in a suitable solvent to pass through a filter. The insoluble impurities remain on the filter, while the clear liquid passes through and is termed *filtrate*.

Experiment 13.—Take a circular piece of filter-paper, fold it in two and double it again to form a quadrant (Fig. 7). Open it out to form a hollow cone

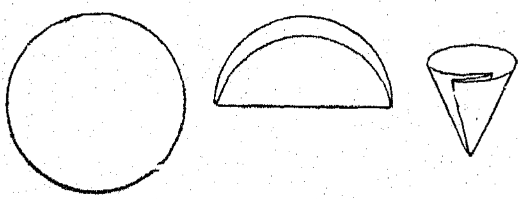


Fig. 7.—Filter-paper.

by taking three folds on one side and one on the other Place it in a glass funnel held in a stand and moisten it with a little water to make it stick to the sides of the funnel. Now take a mixture of equal parts of sugar and powdered charcoal, and treat it with water. Gently pour the liquid on the filter-paper by allowing it to run slowly

down a glass rod, the lower end of which just touches the paper (Fig. 8). The solid particles of charcoal, being unable to pass through the fine pores of the filter-paper, are retained on its surface, while the clear liquid passes through.

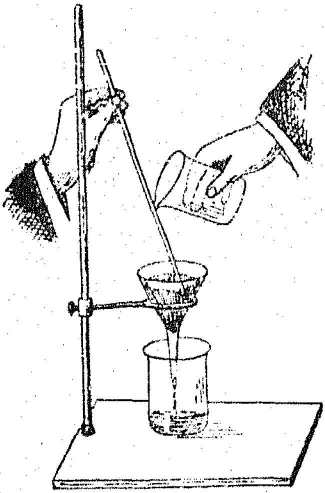


Fig. 8.—Filtration.

The simplest filter, and the one generally employed in the laboratory, is the *filter-paper*, but cloth, charcoal and sand are commonly employed in daily life. In some villages, where pure water is not obtainable for drinking purposes, the muddy water is filtered through charcoal or sand-beds before it is used for drinking. Three earthenware jars (*gharas*) are placed, one above the other, the upper two having holes at the bottom (Fig. 9). The uppermost jar contains muddy water which gently drops into sand contained in the middle one. The sand retains most of the suspended and insoluble impurities, and clear water trickles down into the third vessel.

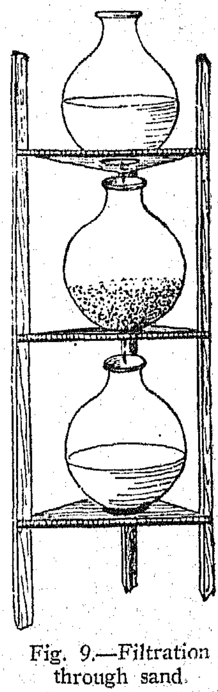


Fig. 9.—Filtration through sand.

In big cities like Lahore or Delhi, water is freed from impurities by filtering it through large filter-beds made up of layers of sand and gravel, or by forcing it under pressure through large filters of unglazed and porous earthenware.

The sugar solution (*shira*) required in the preparation of pudding (*halwa*) is filtered through cloth in order to obtain a clear solution.

Crystallisation.—A hot saturated solution generally contains more of the dissolved substance than a cold one, and therefore when the hot solution is cooled, some of the dissolved substance separates out in definite geometrical forms known as *crystals*. The process by which the solid thus separates out is termed **crystallisation.** In experiment 9, nitre separates out in white crystals.

A solid may be obtained in the crystalline form by evaporating its dilute solution over a gentle heat, or by allowing its hot saturated solution to cool. The slower the evaporation or the cooling, the larger are the crystals obtained.

Experiment 14.—Dissolve some alum (*phitkari*) in water and gently heat the solution in a porcelain dish. Water is driven off and the solution becomes saturated after a time. Allow it to cool; crystals of alum separate out.

Crystallization is used as a means of purification and separation of solids. The impure solid is dissolved in water and the solution is gently heated till it becomes saturated. The hot saturated solution is then allowed to cool. The pure solid crystallises out, while the impurities remain behind in the liquid which is called the *mother liquor.*

This process is employed not only for purifying solids but also for separating two or more

substances which are soluble in the same solvent. Their separation depends upon the facts that (*i*) a solid is generally more soluble in a liquid when hot than when cold, (*ii*) different solids have different solubilities in a liquid at a definite temperature. For instance, nitre (*shora*) is much more soluble in hot than in cold water, whereas the solubility of common salt increases only slightly with the rise of temperature. If, therefore, a hot saturated solution of nitre mixed with common salt be allowed to cool, a large quantity of nitre will crystallise out, while most of the common salt will remain in the solution. If these crystals of nitre, containing a little salt, be again dissolved in water and re-crystallised, the nitre obtained will be purer than before. On repeating the process nitre crystals will be practically free from the whole of common salt. Since separation of solids in such cases is carried on in different fractions, the process is known as *fractional crystallisation*.

11. **Distillation.**—The process of converting a

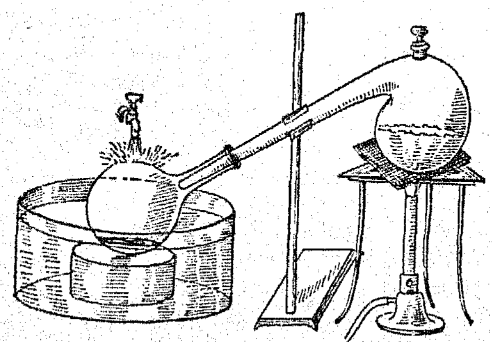


Fig 10.—Distillation.

liquid into its vapour, and condensing and collecting the vapour in a different vessel is called **distillation**. A simple form of apparatus employed for carrying out the process consists of a glass retort *R* fitted to a flask *F* known as *receiver* and placed in a trough of water (Fig. 10). The liquid to be distilled is carefully poured into the retort and heated. When the liquid boils, its vapour passes through the neck of the retort, condenses and collects in the receiver which is kept cool by a regular stream of water.

In order to condense the maximum amount of vapour of a liquid, a more elaborate apparatus, as shown in Fig. 11, is generally used in the laboratory. *F* is a flask fitted with a cork and a delivery tube connected with a Liebig's condenser *C*. The condenser consists of two glass tubes fitting one inside the other. Through the outer tube a slow stream of cold water is allowed to flow by connecting the lower side tube by an india-rubber tubing

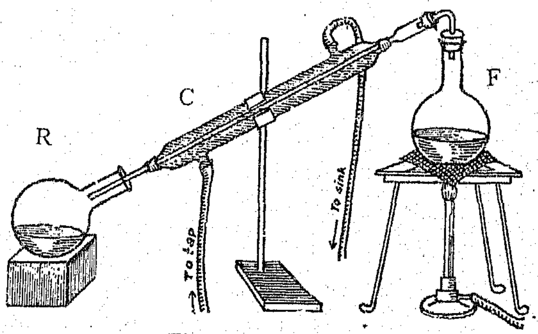


Fig. 11.—Distillation.

with a water-tap, while the upper tube goes to a sink. When the liquid in the flask *F* boils, its

vapour passes through the inner tube of the condenser *C*, condenses and collects in the receiver *R*, leaving behind the dissolved and suspended impurities in the flask.

The process of distillation has a number of practical applications. It is very useful to the chemist in separating liquids from solid and liquid impurities, and to the druggist in obtaining extracts of flowers and plants. The apparatus commonly employed by an Indian druggist in the preparation of extracts (*'araq*) is given in Fig. 12. Distillation is also used to separate two liquids, such as water and alcohol, from each other. Alcohol has a lower boiling point (78° C.) than water, and hence on distillation it passes over first and condenses in the receiver. Distillation is generally used to obtain pure water for drinking purposes from sea-water and other waters not fit for drinking, and to prepare distilled water so often used in the laboratory.

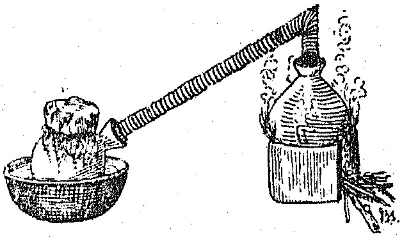


Fig. 12.—The druggist's still.

12. **Separation of mixtures.**—The processes described above are commonly used in separating mixtures of two or more solids with different solubilities. In the case of a mixture of two solids, the mixture is treated with a suitable solvent to dissolve out the soluble constituent, and the solution is filtered. The insoluble residue remains on the filter-paper, while the clear solution

passes through. The soluble substance is then separated from the solvent by evaporation or distillation.

Experiment 15.—Take a little of gunpowder—a mixture of nitre (soluble in water), sulphur (soluble in carbon bisulphide), and charcoal (insoluble in both the solvents)—in a beaker. Treat it with warm water, stir the solution and filter. Wash the residue with water and collect the filtered solution in a dish. Evaporate the solution to dryness: nitre is obtained. Treat the residue consisting of sulphur and charcoal with carbon bisulphide, and filter again. Expose the filtrate to air; carbon bisulphide evaporates and crystals of sulphur are left behind. Dry the charcoal left on the filter-paper by heat or exposure. The three constituents of gunpowder are thus separated.

13. Sublimation.—There are some substances which, when heated, pass straight from the solid to the vaporous state without melting, and the vapour, when cooled, passes back directly into the solid state. They are said to *sublime* and this process of direct passage from the solid to the gaseous state and *vice versa*, is called **sublimation.** Sal-ammoniac (*noshadar*) and camphor behave in this way, and when heated in a test-tube they form a white deposit, known as *sublimate*, on the cooler parts of the tube. The process is used in separating substances that can sublime from those that do not sublime. For instance, in case of arsenic poisoning, white arsenic (*sankhiya*) can be easily detected and separated

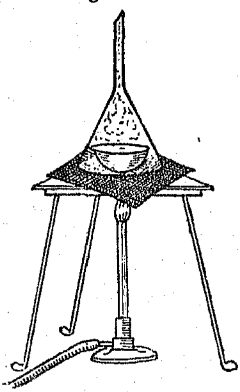


Fig. 13.—Sublimation.

from other substances by this method.

Experiment 16.—Take a mixture of sand and sal-ammoniac in equal parts, and heat it in a dish covered with an inverted glass funnel (Fig. 13). Dense white fumes of sal-ammoniac are formed and condense on the inside of the funnel as a white sublimate, whereas the sand remains behind.

QUESTIONS.

1. How will you find whether the following substances are soluble or insoluble in water:—(1) sand, (2) sulphur, (3) powdered charcoal, (4) lime, and (5) nitre?

2. You are given a mixture of sand, sulphur and nitre. How will you separate the mixture into its constituents and obtain a pure and dry specimen of each?

3. Explain the terms: *solubility, solute, crystallisation, sublimate.* Give examples.

4. You are supplied with some river-water. How will you show that it contains both suspended and dissolved impurities? How will you obtain pure water from it?

5. Describe experiments to show that (*i*) nitre is more soluble in hot than in cold water, (*ii*) lime is less soluble in hot than in cold water.

6. Explain clearly what is meant by a 'crystal.' What are the advantages of forming crystals? How can you obtain good and large crystals of alum? [P. U. Matric., 1919.]

7. How will you proceed to determine the solubility of common salt at the room temperature?

8. Distinguish between the process of evaporation and distillation and give examples of their use.

9. How can a liquid be freed from a solid (*i*) suspended in it, and (*ii*) dissolved in it?

10. You are provided with a mixture of camphor and common salt. How will you separate the mixture into its constituents?

11. You are given a mixture of sand, sulphur and common salt. How would you proceed to recover the various constituents from it? Sketch the apparatus. [P. U. Matric., 1922.]

12. You are given a mixture of powdered glass, ammonium chloride and copper sulphate. Describe in detail the method that you would adopt for separating the constituents. [P. U. Matric., 1923.]

CHAPTER III.

ELEMENTS AND COMPOUNDS.

14. **Elements and compounds.**—When we study different substances, we find that they may be divided into two classes: (1) those which cannot be resolved into anything essentially different from themselves are called *elements*, (2) those which can be split up into two or more dissimilar substances are known as *compounds*. For example, sulphur cannot be made to yield any other kind of matter, nor can iron be resolved into anything else. Sulphur and iron are therefore elements. On the other hand, water is a compound, as it can be broken up into two different substances: oxygen and hydrogen. For the same reason, carbonic acid gas, which can be decomposed into carbon and oxygen, is a compound. Thus, an **element** is a substance which has not so far been split up into simpler substances, and therefore contains, so far as we know, only one kind of matter. A **compound** is a substance which can be broken up into two or more simpler substances.

While the number of compounds is extremely large and is ever increasing, the elements known are limited in number. With the advance of chemical knowledge certain substances, at one time supposed to be elements, have now been discovered to be compounds. Water, for instance, was considered to be an element till 1781, when Cavendish, an English chemist, showed it to be a compound

of oxygen and hydrogen. It is just possible that certain substances, supposed to be elements at present, may turn out to be compounds when subjected to more careful and refined modes of examination. There are at present about ninety elements known. The following is a list of the common and important elements:—

Aluminium	Magnesium
Antimony	Mercury
Arsenic	Nitrogen
Bromine	Oxygen
Calcium	Phosphorus
Carbon	Platinum
Chlorine	Potassium
Copper	Silicon
Gold	Silver
Hydrogen	Sodium
Iodine	Sulphur
Iron	Tin
Lead	Zinc

15. **Mechanical mixture.**—When two or more elements or compounds are brought together, one of the two things happens: either they merely mix together without losing their individual properties, or a chemical change takes place, resulting in the formation of new substances having entirely different properties. In the former case, the result is said to be a **mechanical mixture,** while in the latter it is the formation of a **chemical compound.**

Experiment 17.—Take 7 parts of iron filings and 4 parts of sulphur, and grind the two in a mortar; a greyish powder is obtained. Examine the powder through a pocket lens. Separate particles of iron and sulphur can be easily detected. Now insert a magnet in the powder.

Iron filings will stick to the magnet, and may be separated from sulphur. Here, then we are dealing with a mechanical mixture of iron and sulphur, in which the two constituents retain their individual properties and can be easily separated.

Now heat the mixture in a test-tube; the contents of the tube soon begin to glow. Remove the tube from the flame and allow it to cool. A dark-coloured mass is obtained, in which neither iron nor sulphur can be easily detected, nor can the constituents be separated by any simple means.

If the iron and sulphur in the above experiment be taken in any other proportion, the excess will remain unchanged, and can be easily separated.

A mechanical mixture may be distinguished from a chemical compound in the following ways:—

1. The constituents of a mixture may be present in any ratio, while the constituents of a compound are always combined in a definite proportion. Sugar and sand may be mixed in any ratio, but water is always composed of hydrogen and oxygen in the ratio of one to eight by weight.

2. The constituents of a mixture can usually be separated by simple mechanical means (see Chapter II), while the constituents of a compound cannot be easily separated. A mixture of common salt and sand can be separated into its constituents by treating the mixture with water and filtering the solution. Common salt, on the other hand, cannot be easily separated into its constituents; sodium and chlorine.

3. The properties of a mixture are midway between those of the substances forming the mixture, while the properties of a compound are en-

tirely different from those of the elements present in the compound. The colour of a mixture of equal parts of a white and a black powder, such as sugar and charcoal, will be grey; but mercury oxide, a compound of mercury and oxygen, is red. The properties of air, a mixture of oxygen and nitrogen, are intermediate between those of the two gases; whereas water, a compound of hydrogen and oxygen, has its own distinctive properties.

4. The composition and properties of a mixture may be different in different parts, but the properties of a compound are exactly the same throughout. The proportion of oxygen and nitrogen present in air varies, though slightly, at different places (see Chapter IV). But the composition of water, one part by weight of hydrogen and eight parts of oxygen, is exactly the same everywhere.

The chief points of difference between a mechanical mixture and a chemical compound are summarised in the following table:—

No.	Mechanical Mixture.	Chemical Compound.
1	Constituents mixed in any ratio.	Constituents always present in a definite proportion.
2	Constituents usually separated by simple mechanical means.	Constituents not easily separated.
3	Properties midway between those of constituents.	Properties quite different from those of constituents.
4	Composition and properties different in different parts.	Composition and properties exactly the same throughout.

16. **Solution and a mixture.**—A solution of a solid in a liquid possesses some of the properties

of both the solute and the solvent. Moreover, the solid dissolved in a solvent can be easily recovered from the solution by simple chemical operations, such as evaporation, distillation, or chystallisation. Hence a solution may be considered a mechanical mixture and not a chemical compound. But the behaviour of solutions is rather peculiar and differs from that of mixture in the following respects:—

1. A solution is always a clear and transparent liquid in which the solute is so thoroughly mixed that it cannot be distinguished from the solvent even with the help of a microscope. Particles of common salt dissolved in water cannot be distinguished from those of water. But an ordinary mixture may be opaque and the constituents distinctly visible with the naked eye or a magnifying lens, as in the case of iron filings and powdered sulphur in Experiment 17.

2. A solution is always, like a chemical compound, homogeneous, *i.e.*, every part has the same composition and properties as any other part. But a mixture may or may not be homogeneous. A solution of sugar in water is homogeneous, but sand and sugar, unless very thoroughly mixed, will not form a homogeneous mixture.

3. The composition of a solution can be changed only within certain limits which depend upon the solubility of the dissolved substance, but the composition of a mixture can be varied at will. For instance, the solubility of common salt in water is 35·8 at 30°C., and the amount of the salt dissolved in 100 grams of water can vary between 0 and 35·8 grams at this temperature. This solution will not contain more than 35·8 grams of the salt

for every 100 grams of water at 30° C. But common salt and sugar or sand can be mixed in any ratio.

The student should remember that a solution in the above cases means the solution of a solid in water. In certain cases, solution may be accompanied by chemical changes. Thus, when copper or silver is dissolved in nitric acid, a chemical change takes place and copper nitrate or silver nitrate is formed.

17. **Chemical action.**—When two or more substances, elements or compounds, combine together to form a new substance or a compound decomposes into simpler substances, **chemical action** is said to have taken place. For example, mercury and iodine when rubbed together form iodide of mercury.

Experiment 18.—Rub some mercury with excess of iodine in a mortar. A green iodide of mercury is formed.

Experiment 19.—Heat some nitrate of lead (a white crystalline substance) in a test-tube. The substance decomposes, reddish-brown vapours of nitrogen peroxide are evolved, and a yellow substance (lead oxide) is left behind.

Experiment 20.—Heat some red oxide of mercury in a hard glass test-tube; it decomposes into its constituents: mercury and oxygen. Oxygen escapes and mercury is left behind.

Chemical changes are of two kinds:—

1. **Chemical combination** which consists in the union of two or more substances resulting in the formation of a new compound. For example, iron and sulphur, when heated together, form iron sulphide. In the same way, lead and mercury, when heated in air, take up oxygen and form the corresponding oxides. Experiment 18, in which mer-

cury, when rubbed with iodine, forms a green iodide of mercury, is another illustration of chemical combination.

2. **Chemical decomposition** in which a compound decomposes into two or more simpler substances — elements or compounds. For instance, limestone (calcium carbonate) when strongly heated decomposes into lime (calcium oxide) and carbon dioxide—

Calcium carbonate=calcium oxide+carbon dioxide.

Red oxide of mercury decomposes on heating into its constituents: mercury and oxygen.

In order that chemical action may take place, the substances must be placed in contact with each other. The closer the contact, the quicker the chemical combination.

Experiment 21.—Heat some granulated zinc with a lump of sulphur; little or no action takes place. Now heat a mixture of zinc dust and flowers of sulphur. A vigorous action due to a close contact occurs, and a bright glow of light is observed.

The best way to obtain close contact between two substances is to turn one or both of them into the form of a solution.

Experiment 22.—Mix some dry and powdered lead nitrate and potassium iodide together in a beaker; no change occurs. Now add water to the mixture; an orange-red substance is formed. The contact is so close that the change appears to be instantaneous.

Experiment 23.—Place a mixture of dry washing soda and tartaric acid (*tartari*) in a glass tumbler and add water. A brisk effervescence takes place.

Chemical action is aided by heat, light, pressure and several other agencies. A mixture of iron and sulphur remains unchanged unless heat is applied

to it. A mixture of hydrogen and chlorine combines vigorously in the presence of sunlight. Gunpowder will explode when subjected to pressure.

Experiment 24.—Powder some potassium chlorate (*potash*) and antimony sulphide (*mansal*) in *separate* mortars. Take a small quantity of each and mix the two with a feather, avoiding all unnecessary pressure. Place the mixture on the floor and strike it with a hammer. An explosive noise is produced.

When chemical action takes place, heat, light, or electricity is generally produced.

Experiment 25.—Place a lump of burnt lime on a stone slab and throw some water on it. The lime becomes very hot and a hissing noise is produced.

Experiment 26.—Place one or two pieces of dry phosphorus on an iron plate, and sprinkle a little of iodine over it. Phosphorus and iodine combine together and ignite spontaneously.

Note.—Phosphorus must be handled very carefully, as it catches fire spontaneously when exposed to air.

The generation of electric current in a simple voltaic cell is due to the action of dilute sulphuric acid on zinc.

18. **Chemical affinity.**—Some elements combine with one another more readily than others. They possess, so to say, a peculiar attraction for one another. This attraction which certain elements have for others is known as **chemical affinity.** For example, phosphorus when exposed to air combines with the oxygen of air so readily that the heat produced is sufficient to ignite it. Sodium readily reacts with water at the ordinary temperature, combining with oxygen and giving off hydrogen. Silver and gold, on the other hand, may be left exposed to air for any length of time

without suffering any change. Phosphorus and sodium are said to have a strong affinity for oxygen, while silver and gold have very little affinity for it. In the same way, chlorine has a very great affinity for hydrogen, because it readily combines with the latter when the two are mixed together and the mixture placed in sunlight. Chlorine has very little tendency to unite with oxygen and is therefore said to have little affinity for it.

The exact nature of chemical affinity is unknown. We do not know why certain elements have a stronger affinity for one another than others. Chemical affinity is comparable to attraction of gravity. It is to the chemist what gravity is to the physicist. But whereas gravity acts at all distances, affinity can operate only at extremely small distances.

QUESTIONS.

1. What do you understand by the terms: *elements* and *compounds?* Give reasons to prove that sulphur is an element and not a chemical compound.
2. Classify the following substances into elements and compounds:—steam, iron filings, common salt, sugar, diamond, alum, zinc dust and nitre.
3. What do you understand by an *element*, a *mixture* and a *compound?* How would you classify the following:—diamond, ice-cream, chalk, sulphur, gunpowder and brass! Give reasons for your answers. [P. U. Matric., 1923.]
4. How do chemical compounds differ from mechanical mixtures? Give examples to illustrate your answer.
5. You are supplied with a powder containing copper and sulphur. How will you proceed to determine whether it is a mixture or a compound of the two elements?
6. What is meant by *chemical action?* Give examples.
7. What are the conditions necessary for bringing about a chemical action?

8. Describe experiments to show the action of (*i*) heat, (*ii*) light, and (*iii*) electricity in causing chemical changes.

9. Name the agencies that promote chemical action. Give examples.

10. What do you understand by *chemical combination* and *chemical decomposition?* Give examples.

11. Explain, as clearly as you can, the meaning of the term *chemical affinity*.

CHAPTER IV.

AIR.

19. **The air.**—The atmosphere, commonly called the *air*, is the gaseous envelope which surrounds the earth. Its presence, though easily felt when air is set in rapid motion, may be proved in a number of ways. On waving a sheet of cardboard in air, we feel that something offers resistance to its motion. When we move a fan, we feel a current of air striking our body. Its existence in an apparently empty vessel may be easily proved by the following simple experiment:—

Experiment 27.—Invert a glass tumbler and press it on the surface of water contained in a trough. The water does not rise and fill the tumbler, because the air which is present in the tumbler offers resistance to the water and does not allow it to rise in the tumbler.

Air is a form of matter and occupies space and possesses weight. It is difficult to form an exact idea of the height to which the air extends, but it is found to have an appreciable existence up to a height of about 200 miles above the surface of the earth. Its weight may be found by taking a flask fitted with a stop-cock and weighing it before and after exhausting the air present in it. The weight of one litre of air is found to be about 1·293 grams at the ordinary temperature and pressure. The weight of air in a room of ordinary dimensions (15 feet $\times$ 14 feet $\times$ 12 feet) will be about 200 pounds, or nearly 2 maunds and 17 seers.

20. **Composition of air.**—Formerly air was regarded as one of the four so-called elements: air, water, fire and earth. But now we know that air is neither an element nor a compound, but a mixture of several gases, the two chief constituents being oxygen and nitrogen. Air also contains a small proportion of carbon dioxide, water vapour, and negligible quantities of ammonia, oxides of nitrogen, and dust particles.

Nitrogen and oxygen are found to be present in the air in the ratio of 79 to 21 by volume and 77 to 23 by weight. The relative proportions of the two gases by volume can be determined by removing the oxygen from a known volume of air by a chemical action, such as the burning of phosphorus or candle, or the rusting of iron.

Experiment 28.—Place a small porcelain dish on a thin and broad piece of wood or cork, floating on water contained in a trough (Fig. 14). Cover the dish with a bell-jar, and allow the jar to rest on two glass rods placed at the bottom of the trough. Remove the stopper of the jar and mark the level of the water in it by a strip of gummed paper. Divide the space above the water level into five equal portions. Now remove the jar and place a small piece of dry phosphorus in the dish. Carefully replace the bell-jar and ignite the phosphorus by touching it with a heated end of an iron rod. Quickly replace the stopper. The phosphorus burns and gives out dense white fumes of oxide of phosphorus. The fumes disappear on cooling, and the water rises in the jar and

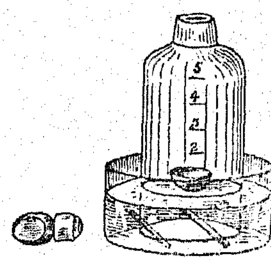


Fig. 14.—Composition of air by burning phosphorus.

fills one-fifth of the space originally occupied by the air.

Remove the stopper of the bell-jar again and introduce a burning taper in the jar. The flame is at once extinguished, showing that the remaining gas is not the same as before.

Repeat the above experiment with a burning candle but replace the water in the trough by caustic soda solution to absorb carbon dioxide produced by the burning of the candle. The candle takes up oxygen of the air and forms carbon dioxide which is absorbed by caustic soda solution. The water formed condenses to the liquid state and the water rises in the jar to fill about one-fifth of the space originally occupied by the air.

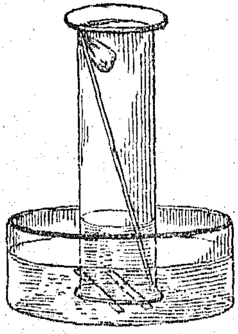


Fig. 15.—Rusting of iron.

Experiment 29.—Take some iron filings in a muslin bag and tie the bag to one end of a glass rod. Moisten the bag and place it in a wide-mouthed bottle or a jar inverted over water in a trough (Fig. 15). Allow the apparatus to stand for a few days. The water will rise in the jar and fill one-fifth of the space previously occupied by the air. Remove the bag and examine the filings; they will be found covered with iron rust (*zang*).

From the above experiments we learn that air is composed of two gases: (1) that which helps substances to burn in it. It forms the *active part* of air, and is known as **oxygen**; (2) that which does not allow substances to burn in it and is the *inactive part* of air; this is known as **nitrogen.** About one-fifth of air is oxygen, and four-fifths is nitrogen. When phosphorus or a candle burns in an enclosed space of air, or when iron is allowed to rust, action goes on until one-fifth of the air

has disappeared, and then the action stops. This shows that in each case it is the same constituent, *viz.*, oxygen that combines with the burning substances.

21. **Air a mixture.**—There are several reasons to prove that air is a mixture and not a chemical compound:—

1. The properties of air are midway between those of oxygen and nitrogen. In oxygen substances burn vigorously, nitrogen does not allow them to burn at all, while air stands between the two extremes and is a moderate supporter of combustion.

2. When oxygen and nitrogen are mixed in the proportion in which they exist in air (*i.e.* 21 : 79), no heat is produced and no change in volume takes place, but the mixture obtained shows all the properties of air.

3. The proportions in which oxygen and nitrogen are present in air vary a little at different times and places, but the constituents of a compound are always present in a fixed ratio. Air at all places does not contain oxygen and nitrogen in the ratio of 21 : 79 by volume. But water is always found to consist of 1 part of oxygen and 2 parts of hydrogen by volume.

4. The constituents of air can be easily separated by simple means, such as solution and liquefaction. Nitrogen is about half as soluble as oxygen in water, and therefore when air is shaken with water, more of oxygen than nitrogen dissolves. If the water be now boiled, the expelled gas is found to be richer in oxygen than the original air. Ordinary air consists of about 1 part of

oxygen and 4 parts of nitrogen, whereas the air expelled from boiling water contains about 1 part of oxygen and 2 parts of nitrogen. Had air been a chemical compound, it would have dissolved in water and been expelled from it as a whole.

When air is liquefied and the temperature of the liquid air allowed to rise slowly, a large proportion of nitrogen evaporates first, leaving the remaining liquid richer in oxygen.

22. **Other gases in air.**—Oxygen and nitrogen, as we have seen, are the main constituents of air. Besides these, small and varying quantities of moisture and carbon dioxide, and minute traces of ozone, ammonia, dust particles and some rare gases are also present in air. The average proportions in which these are found in air are given in the table below.

AVERAGE COMPOSITION OF AIR.

No.	Name of the constituent.	Percentage Composition.	
		By volume.	By weight.
1.	Nitrogen	78.0	75.5
2.	Oxygen	20·9	23.1
3.	Water vapour	Variable	Variable
4.	Carbon dioxide	0·03	0.04
5.	Ozone, ammonia and other gases ...	0·93	1.3

Water vapour. In rainy weather, common salt generally becomes wet, because it takes up moisture from the atmosphere. When a vessel contain-

ing ice or ice-cold water is left exposed to air, moisture is seen deposited in the form of minute water drops on the outer surface of the vessel.

Experiment 30.—Take a glass tumbler, clean it thoroughly, and wipe off the outside such that no moisture is left on the outer surface. Put some pieces of ice in the tumbler and let it stand for a while in a room. Water vapour present in the air of the room gets condensed, and small drops of water are observed on the outside surface of the tumbler.

Calcium chloride and caustic soda become liquid on exposure due to the moisture absorbed from air. Such substances that absorb moisture when exposed to air are termed **hygroscopic.** The presence of moisture in the atmosphere is mainly due to the evaporation ot water that is continuously going on from the surface of lakes, rivers, seas, etc.

Carbon dioxide.—The amount of carbon dioxide varies in different localities, though not through such a wide range as the water vapour. It is present in larger quantities in closely inhabited towns than in the open country. Under ordinary conditions its proportion varies between 3 and 6 parts in 10,000 of air, but in badly-ventilated rooms it may go up to ten times the amount. Its presence in air is due to the breathing of animals, burning of wood, coal, etc., and the decay of animal and vegetable matter. The existence of carbon dioxide in air may be shown by the following experiment:—

Experiment 31.—Treat some quicklime with excess of water. A milky liquid is obtained. Filter the liquid. The clear filtrate is known as *lime-water*. Place the lime-water in a shallow dish and expose it to air. A white

crust of chalk (calcium carbonate) is formed on the surface, showing the presence of carbon dioxide in air.

Ozone.—Ozone is present in small traces in air. However, it is found in comparatively large quantities in the air above the sea or near the sea-coast, as it is said to be formed during the evaporation of water. The health-giving properties of the sea-air are due to the presence of ozone.

QUESTIONS.

1. Describe an experiment to show that air occupies space.
2. Describe an experiment to prove that air is composed of at least two gases. Give the properties of each gas.
3. How would you show that only a portion of air takes part in the burning of substances? Draw a diagram of the apparatus used.
4. What are the main constituents of air? In what proportion are they present (*i*) by volume, and (*ii*) by weight?
5. How would you proceed to show that air contains about four-fifths of its volume of nitrogen?
6. A piece of phosphorus is burnt in a stoppered bottle containing 150 c.c. of air. The bottle is then inverted under water and the stopper removed. What volume of water enters the bottle? What do you learn from the experiment?
7. Justify the statement: 'Air is a mixture and not a chemical compound.'
8. How would you prove the presence of water vapour and carbon dioxide in air? How would you account for their presence in air?

CHAPTER V.

NITROGEN AND OXYGEN.

Nitrogen.

23. **Nitrogen** is an inert gas and forms about four-fifths of air. Unlike oxygen, it does not combine directly with most of the elements, and can therefore be conveniently obtained by treating a limited quantity of air with a substance which combines easily with oxygen but has no action on nitrogen. The substances commonly chosen for the purpose are phosphorus (expt. 28, page 30) and copper. The gas obtained by this method is, however, never pure.

Experiment 32.—Take a hard glass tube and place some copper turnings in it. Connect one end of the tube with a big bottle fitted as shown in Fig. 16, and the other with a delivery tube passing under a test-tube. Fill the test-tube completely with water, and invert it over a bee-hive shelf placed in a trough of water. Heat the copper to redness, and allow a slow

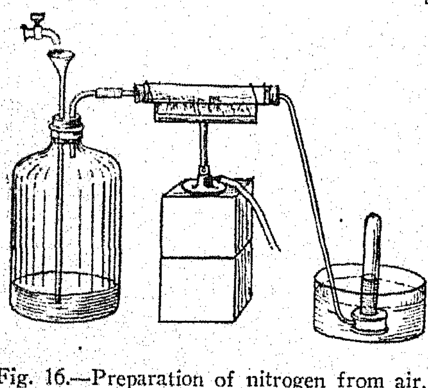


Fig. 16.—Preparation of nitrogen from air.

and constant stream of air to pass over it by running water in the bottle from a water-tap. The oxygen of the air will be taken up by the heated copper, and nitrogen collected in the test-tube—

copper + air = copper oxide + nitrogen.

Remove the tube when full, and introduce a burning taper in it. The taper is extinguished, showing that nitrogen does not support burning.

Nitrogen is easily prepared in the laboratory by heating ammonium nitrite which decomposes to give nitrogen and water—

Ammonium nitrite = nitrogen + water.

Experiment 33.—Take some dry ammonium nitrite in a hard glass test-tube fitted with a wide delivery tube dipping under water in a trough (Fig. 17). Heat the test-tube; the substance melts and then decomposes to give nitrogen. Collect the gas in a jar filled with water and inverted over a bee-hive shelf in the trough.

24. **Laboratory and commercial methods.**—The above methods for the preparation of nitrogen are known as **laboratory methods,** while those employed for preparing substances on a large scale are called **commercial methods.** In the laboratory we prefer methods which are easy and instructive. The cost involved in the preparation of a substance by these methods does not matter much. In commerce, on the other hand, economy is the determining factor, and

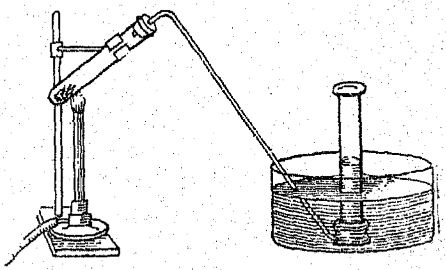


Fig. 17.—Preparation of nitrogen.

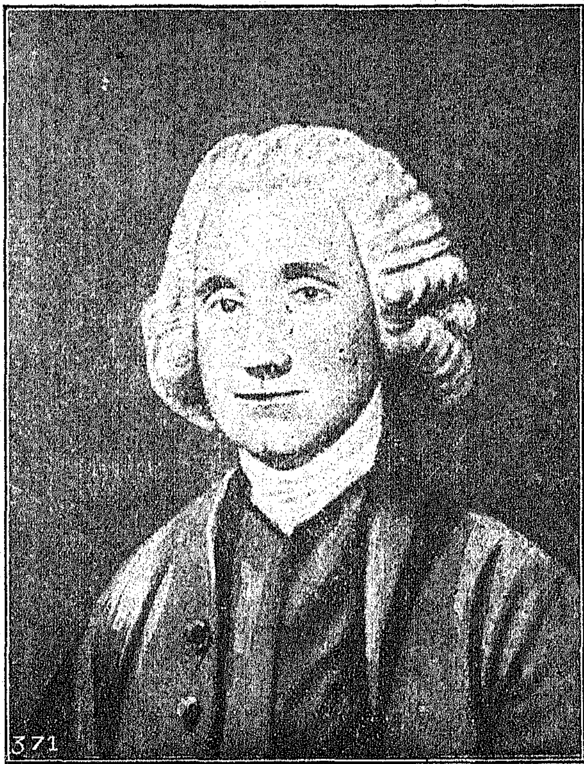
methods which are comparatively cheap are always preferred. Moreover, the value of by-products (*i.e.* substances obtained in addition to the one required) in the manufacture of substances is of considerable importance for the successful working of a commercial method.

Nitrogen is prepared on a commercial scale from air which has been liquefied by the combined effect of high pressure and low temperature. Liquid air is mainly a mixture of liquid oxygen and liquid nitrogen. The boiling point of the latter is lower than that of the former. Hence when liquid air is allowed to evaporate, nitrogen changes to vapour first, and is thus separated from oxygen which is left behind.

25. Nitrogen is a colourless gas without smell or taste. Unlike oxygen, it combines directly with a very few elements, and does not allow substances to burn in it. It is not poisonous, but animals die in it for want of oxygen. It is nearly half as soluble as oxygen in water. Its use in air is to moderate the action of oxygen which, if present alone, would be too vigorous a supporter of life.

Oxygen.

26. **History.**—Priestley, an English chemist, was particularly fond of studying the nature of gases evolved from certain substances when heated. Once in 1774, he happened to expose red oxide of mercury to the rays of the sun concentrated by a large burning glass, and found to his surprise that the red oxide turned black and a gas was given off. He then found that substances burn much more brilliantly in this gas than in air. Thus was



Joseph Priestley (1732—1804)

(*To face page* 38).

oxygen discovered.

27. **Occurrence.**—Oxygen is the most widely distributed of all the elements in nature. It occurs both in the free state and in combination with other elements. In the free state, it is found in air of which it forms nearly one-fifth by volume. In combination, it constitutes eight-ninths of the weight of water, and about one-half the total weight of the earth's crust. It also forms an important constituent of all animal and vegetable substances.

28. **Preparation.**—(1). Oxygen is prepared on a small scale by heating red oxide of mercury, when oxygen is evolved and mercury left behind—

Mercury oxide = mercury + oxygen.

Experiment 34.—Heat some red oxide of mercury in a hard glass test-tube. It turns black and then decomposes to give out oxygen. Mercury is left behind in the form of a mirror. Introduce a glowing splinter into the test-tube; it bursts into flame.

(2). Oxygen can be readily obtained for laboratory purposes by heating potassium chlorate mixed with black oxide of manganese. The potassium chlorate decomposes into potassium chloride and oxygen. The presence of manganese dioxide causes the decomposition to take place at a much lower temperature than when the chlorate is heated alone. Manganese dioxide can be recovered unchanged at the end of the experiment, and therefore plays no apparent part in the reaction. Such a substance, that by its mere presence makes a chemical action go much faster and itself remains chemically unchanged at the end of the action, is said to be a **catalytic agent.** Its action may be compared to that of a whip to a sluggish horse.

or of an oil to a machine.—

Potassium chlorate+(manganese dioxide)
=potassium chloride+oxygen+(manganese dioxide).

Experiment 35.—Mix some powdered potassium chlorate with about one-fourth of its weight of manganese dioxide. Place the mixture in a hard glass test-tube fitted with a delivery-tube leading to an inverted jar of water in a pneumatic trough. Fix the test-tube in an iron clamp in an inclined position, as shown in Fig. 18. Heat the mixture, and collect four or five jars of the gas. Keep each of them covered with a glass disc, and test the properties of the gas (see below).

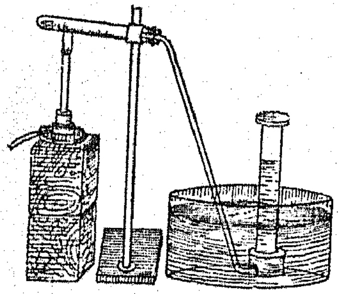


Fig. 18.—Preparation of oxygen.

When all the jars have been filled with the gas, remove the delivery tube *first* and *then* the flame, otherwise the water will rush back into the hard glass tube and crack it.

(3). On a commercial scale, oxygen is prepared either from liquid air or by electrolysis of water.

(*i*). *From liquid air.*—Most of the oxygen is now manufactured by this method. Here advantage is taken of the fact that liquid nitrogen boils at a lower temperature than oxygen. When air is liquefied by intense cold and pressure and the liquid air allowed to evaporate, nitrogen escapes first, and the remaining liquid becomes richer in oxygen. By this method oxygen can be

obtained in a fairly pure state.

(*ii*). *By electrolysis of water.*—When water is acidified with a few drops of sulphuric acid and electric current passed through it, it is split up into oxygen and hydrogen which are collected on two separate plates known as *electrodes* (see experiment 53, page 62).

29. **Properties.**—The properties of a substance may be divided into two kinds: (1) *physical*—those which can be detected by the senses, such as colour, smell, taste, weight and solubility; (2) *chemical*—those which can be examined by the behaviour of the substance towards other reagents.

Physical.—Oxygen is a colourless gas without taste or smell. It is slightly soluble in water, and it is due to this dissolved gas that fish and other aquatic animals can live under water. It is a little heavier than air, and about sixteen times as heavy as hydrogen.

Chemical.—At the ordinary temperature, oxygen is not a very active gas, but at high temperatures it can combine with nearly all the other elements. It does not burn itself, but other substances burn in it very brilliantly. Carbon, sulphur and phosphorus burn more brightly in this gas than in air, forming substances which turn blue litmus solution red; while sodium, potassium and calcium form substances which, when dissolved in water, turn red litmus blue.

Experiment 36.—Take a jar of the gas prepared in experiment 35, and introduce a glowing splinter in it. The splinter bursts into flame.

The above property of oxygen is used as a test for the gas.

Experiment 37.—Lower a burning candle in another jar. The candle burns more brightly. Now pour some lime-water in the jar. It turns milky, showing that carbon dioxide is formed when a candle burns in oxygen.

Experiment 38.—Place a little of dry phosphorus in a deflagrating spoon (a metal cup fixed to a stout metallic wire, Fig. 19), and heat on a flame. When phosphorus begins to burn, introduce it in a jar of the gas. The phosphorus burns with a dazzling light and dense white fumes of oxide of phosphorus are produced.

Repeat the experiment with sulphur; it burns with a beautiful bluish flame.

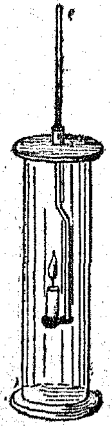


Fig. 19.—Burning of a candle in oxygen.

After the phosphorus or sulphur has burnt, put some litmus solution in the jar and shake. The solution turns red.

Experiment 39.—Ignite a piece of sodium or potassium on a deflagrating spoon and plunge it in another jar of oxygen. The metal burns with an increased brilliancy in the gas. Pour some litmus solution (which has been made red by the addition of a drop of very dilute sulphuric acid) into the jar; the solution turns blue.

NOTE.—*Sodium or potassium must be cut with a dry knife, as it catches fire in contact with moisture.*

Iron, which burns with difficulty in air, burns with a brilliant shower of sparks in oxygen, giving a black solid (iron oxide) which has no action on litmus.

Experiment 40.—Take a few pieces of very thin iron or steel wire and twist them in the form of a small bundle. Fasten this to the end of a deflagrating spoon and heat. Now dip the heated end in a little powdered sulphur so that a small quantity of sulphur adheres to it. Heat the wire again to ignite the sulphur and introduce it in a jar of oxygen. The sulphur burns brightly, and the heat produced causes the iron

wire to burn with brilliant sparks (Fig. 20). Small molten globules of iron oxide fall to the bottom of the jar. It is advisable to place some sand or water at the bottom of the jar before introducing the iron wire, as otherwise the jar might be broken by the hot globules. Pour some litmus solution in the jar; the colour remains unchanged.

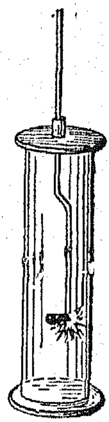


Fig 20. Burning of iron in oxygen.

30. **Oxidation.**—The act of combination of an element with oxygen is called **oxidation,** and like other chemical actions it is generally attended with the production of heat. But when the chemical combination is so rapid that both heat and light are produced, the phenomenon is known as **combustion.** The substance which burns is said to be *combustible*, and the gas in which it burns is called a *supporter of combustion.* The terms *combustible* and *supporter of combustion* are, however, purely relative, *i.e.* a substance which is combustible in the ordinary sense, may become a supporter of combustion when the conditions are suitably changed. We usually have hydrogen or coal-gas burning in air, and therefore the former is called combustible and the latter a supporter of combustion. But combustion would take place equally well, if the conditions are reversed and a jet of air made to issue in an atmosphere of hydrogen or coal-gas. In this case, then, air would be called combustible and hydrogen or coal-gas a supporter of combustion.

The compounds formed by the union of elements with oxygen are termed **oxides.** Oxides may be divided into three main classes:—

(1). **Acidic oxides,** which combine with water to form *acids* which turn blue litmus red. For example, the oxides of carbon, sulphur and phosphorus. These oxides combine with water and form carbonic acid, sulphuric acid and phosphoric acid respectively.

(2). **Basic oxides,** which react with acids to form salt and water. Oxides of sodium, potassium and magnesium are examples of basic oxides. Some of these oxides, such as those of sodium, potassium and calcium, are very soluble in water and form alkalies. These alkalies are hydroxides of the metals, and when dissolved in water possess a soapy touch and turn red litmus blue.

(3). **Neutral oxides.**—These are neutral to litmus. Oxides of hydrogen (water) and carbon (carbon monoxide) are familiar examples.

31. **Uses of oxygen.**—The atmospheric oxygen is indispensable to human life. 'Compressed oxygen,' stored in steel cylinders under a high pressure is used for artificial respiration in hospitals. In the recent Mount Everest expedition (1922), the explorers had to carry 'compressed oxygen' for artificial breathing at great altitudes where air is so much rarefied. Mixed with hydrogen, it is used in the 'oxy-hydrogen' flame for producing *lime-light* and for fusing metals like platinum and steel.

QUESTIONS.

1. Give a short account of the preparation and properties of nitrogen.

2. How was oxygen discovered? How is the gas prepared in the laboratory? Sketch the apparatus used.

3. In what respects does oxygen resemble and differ from

air? Explain by examples.

4. What is meant by the terms: *combustible* and *supporter of combustion, oxidation* and *reduction?* Give examples to illustrate your answer.

5. What changes take place when a candle is burnt in oxygen? How would you show that these changes have taken place?

6. Describe the properties of the chief constituents of the atmosphere. In what proportions are they found?

7. A lighted taper burns more brilliantly in oxygen than in the air. Explain this. [P. U. Matric., 1920.]

8. Name the three kinds of oxides. How do they differ from one another? Give examples of each.

CHAPTER VI.

WATER.

32. **Water** plays such an important part in daily life that every one is familiar with it. It is very widely distributed in nature, the ocean alone covering about three-fourths of the surface of the earth. It was regarded as an element till 1781, when Cavendish, a great English chemist, showed it to be a compound of hydrogen and oxygen combined together in the ratio of two to one by volume.

Water occurs so plentifully in nature that it is unnecessary to prepare it by a chemical process. When required in a *pure* condition, it is obtained by the distillation of ordinary water (page 14).

33. **Physical properties of water.**—Water, when pure, is a tasteless and odourless liquid. It is colourless in small quantities, but shows a greenish-blue tinge when seen through thick layers. The beautiful blue colour of the *Panja Sahib tank* (Hasan Abdal), *Choha Saidan Shah* (Jhelum) and the *Verinag spring* (Kashmir) are a few well-known examples. Like most other substances water can exist in the solid, liquid and gaseous states. It expands when heated, and at 100°C. it is converted into steam which occupies 1,700 times the volume of water. When cooled, water gradually contracts until it reaches 4°C. At this temperature water has the maximum density. On cooling below 4°C. water again expands until at 0°C. it begins to be converted into ice, one volume of water at 0°C.

giving 1·09 volume of ice at the same temperature. Ice is therefore lighter than water and floats on its surface. Water is one of the commonest and most important solvents, and can dissolve a large number of substances.

34. **Natural waters.**—On account of its great solvent power water is never found in an absolutely pure state. Even rain-water, the purest form of natural waters, contains dissolved gases and dust particles.

1. **Rain-water.**—Water is being constantly evaporated from the surface of ponds, lakes, rivers and seas by the heat of the sun. As these enormous quantities of water vapour rise up and condense in the upper regions of the atmosphere, they fall down as *rain*. Rain-water should, therefore, be as pure as distilled water. But on its way to the earth it takes up dust particles, carbon dioxide, ammonia, etc. from the atmosphere. Rain-water collected in towns is much less pure than the one collected in the country, where the atmosphere contains less impurities. The amount of the solid matter present in rain-water is very small at the end of a shower, as the impurities are gradually washed away from the atmosphere.

2. **Spring-water.**—About 25 to 40 per cent. of the rain falling on the surface of the earth soaks into the ground. As this water travels under the ground, it, sooner or later, meets an impervious bed of rock, and being unable to pass through it is forced up to the surface by the pressure of the water at higher level and appears as *spring-water*. It contains more dissolved matter than rain-water; for it remains in contact with the earth and takes

up certain mineral substances from the soil. The nature and amount of the dissolved matter varies with the nature of the soil with which the spring water comes into contact.

If spring water contains a sufficiently large quantity of dissolved matter to impart it a marked taste, it is known as *mineral water* to distinguish it from *fresh water* which contains only a small quantity of dissolved impurities and has no decided taste. Mineral waters are generally named after the important substance dissolved in them, and are used in medicine. *Sulphur waters,* containing compounds of sulphur and obtained from *Sohna* (Rohtak), *Jwalamukhi* (Dharamsala), *Nathyagali* (Abbottabad) and several other sulphur springs are used in curing skin diseases.

3. **River-water.**—Rain and spring-waters collect into streams and rivers. As a river moves on, it dissolves mineral matter from the soil over which it flows, and takes up organic matter from the plants growing on its sides and bottom. It also gets contaminated with the drainage from the towns and villages near which it passes. Since it is constantly moving, it can also carry mud, sand and other solids in the suspended form. Hence river-water is very impure and contains both suspended and dissolved impurities.

Experiment 41.—Take equal quantities (say 50 c.c.) of (*i*) distilled water, (*ii*) tap-water, (*iii*) well-water, (*iv*) tank-water, and (*v*) river-water in separate evaporating dishes. Evaporate the water in each dish to dryness, and notice the difference in the amount of the solid left in each case.

4. **Sea-water.**—Rivers carry their impurities to the sea. Since evaporation is constantly going on

from the surface of the sea, the amount of the solid matter is gradually increasing. Hence it is that in spite of the removal of several impurities by deposition, sea-water is the most impure form of natural waters. In inland seas, where the water carried off by evaporation is greater than that brought in by rivers, the amount of the solid matter is comparatively large. The Dead Sea contains about 23 per cent. of the dissolved matter, while the Mediterranean Sea with its numerous tributaries and less evaporation contains only 3·4 per cent. of solids in solution.

35. **Drinking-water.**—When water is to be used for drinking purposes, it must be colourless, odourless, and free from injurious matter. Sometimes drinking water, such as water of wells and canals, may contain germs of cholera, typhoid fever, etc. derived from the sewage. This water must be boiled to kill the germs before it is used for drinking.

Water to be used for drinking purposes in big towns is stored in big reservoirs where it is at rest and most of the suspended impurities settle down. It is then generally purified by filtration through large filter-beds of sand and gravel. The filtered water is treated with potassium permanganate, or with chlorine or ozone to destroy the disease-producing bacteria that may be present. The well-water used in villages for drinking purposes is often purified by the addition of potassium permanganate (*lal dawai*).

36. **Hard and soft water.**—When we wash our hands or clothes with tap-water, we find that a lather is produced very easily. But in the case

of spring or river-water, a good deal of soap has to be used before the lather appears. Water in which soap lathers easily is known as **soft water,** while that which produces a lather with difficulty is said to be **hard water.**

Experiment 42.—Take some tap-water in a dish and rub your hands with an ordinary soap in it; a lather is soon produced. Repeat the experiment with tank or river-water; no lather appears, but a curd is formed. Go on rubbing soap in water; lather will appear only after a large quantity of soap has been consumed. Tap-water is therefore soft water, while tank or river-water is hard water.

The hardness of water is due to the presence of calcium and magnesium salts in solution. These salts combine with soap and form insoluble compounds, which appear in the form of a curd. So long as these salts are not completely removed, all the soap is used up in forming these insoluble compounds, and no lather appears. But when they are all removed, any further addition of soap at once produces a lather. As a good deal of soap is wasted in the formation of curd which has no cleansing properties, hard water is not economical for washing purposes.

Experiment 43.—Take some tap-water in a dish and rub soap in it, as in the last experiment. A lather is soon produced. Now add a little calcium or magnesium chloride to the same. The lather disappears and a scum is formed on the surface of the water. Again rub the soap in it; a considerable amount of soap will be used before the lather reappears.

Hardness of water is of two kinds: *temporary* and *permanent.* Temporary hardness is due to the presence of calcium and magnesium carbonates in solution. These carbonates are insoluble in water,

but dissolve in water containing carbon dioxide due to the formation of soluble bicarbonates. Natural water generally contains some carbon dioxide dissolved from the air, and when it passes over layers of *lime* and *magnesia*, it dissolves a portion of these carbonates and is rendered hard—

Calcium carbonate+water+carbon dioxide
= calcium bicarbonate.

On boiling, the carbon dioxide is given off, the carbonates settle down, and the water becomes *soft*—

Calcium bicarbonate
(soluble)
= calcium carbonate+carbon dioxide+water
(insoluble)

This explains the formation of a white crust, known as *fur*, on the inside of a kettle in which hard water is boiled. Thus temporary hard water can be *softened* by boiling.

Temporary hardness can also be removed by the addition of lime, which combines with the soluble bicarbonates of calcium and magnesium and converts them into the insoluble carbonates—

Calcium bicarbonate+lime = calcium carbonate.
(soluble) (insoluble)

Permanent hardness of water is due to the presence of chlorides and sulphates of calcium and magnesium, and cannot be removed by boiling. Permanent hard water can be softened by the addition of washing soda (sodium carbonate). The sodium carbonate reacts with the chlorides and sulphates of calcium and magnesium, and converts them into the insoluble carbonates which are pre-

cipitated—

Calcium sulphate + sodium carbonate
(slightly soluble) (soluble)

= calcium carbonate + sodium sulphate.
(insoluble) (soluble)

The soluble sulphate or chloride of sodium, which is formed at the same time, does not affect the cleansing properties of water.

Another method of softening permanent hard water is to distil it, when pure water is obtained as the distillate, while the dissolved solid is left behind.

37. **Determination of hardness.**—Water, as we have seen, may be soft or hard, and the hardness may be temporary or permanent. In order to find out whether a given sample of water is hard or soft, and if it is hard, whether the hardness is temporary or permanent, we treat the given sample of water with soap solution. If on shaking, lather is at once produced and no curd is formed, water is soft. If on the other hand, a curd is produced and no lather is formed, it is hard.

A fresh quantity of this water, if found hard is boiled. If a precipitate is obtained, the hardness is temporary, otherwise it is permanent.

Experiment 44.—Prepare some soap solution by dissolving 1 gram of soap in a mixture of 30 c.c. of distilled water and 15 c.c. of alcohol or rectified spirit.

Take some of the given specimen of water in a test-tube, add a little of the soap solution to it, and shake. If a curdy precipitate is formed, the water is hard, otherwise soft.

If the water is found to be hard, again take a little of it in a test-tube and boil. If a precipitate is formed, the hardness is temporary. Filter. Add the soap solution

to the filtrate, and shake. A lather is produced and no curd is formed. The water is rendered soft by boiling.

If, however, no precipitate is formed on boiling, the water is permanently hard. Add sodium carbonate solution to it and see that a white precipitate is formed. Filter and add the soap solution to the filtrate. The curdy precipitate does not appear, showing that the permanent hardness is removed by the addition of sodium carbonate.

38. **Chemical properties of water.**—As we have already learnt, water has the property of combining with the oxides of carbon, sulphur and phosphorus to form **acids**, and with those of sodium, potassium and calcium to form **bases** (p. 44).

Water has a peculiar property of combining with a number of compounds. Such compounds crystallise out from their aqueous solutions with a definite number of water molecules. Thus, copper sulphate (*nila tutia*) crystallises out with five molecules of water, while iron sulphate (*hira kasis*) with seven molecules. This water is more or less loosely attached to the compound, and is driven out on heating. It is known as the **water of crystallisation** owing to the fact that when it is driven out, the compound generally loses its crystalline form and crumbles to a powder.

The water of crystallisation is held by these compounds with varying degrees of firmness. Some of these compounds give up a part or whole of it simply on exposure to air. Such substances are known as **efflorescent** substances. For example, washing soda (sodium carbonate) and magnesium sulphate lose most of their water of crystallisation on exposure and fall to a powder. Others require strong heating before they lose their water of

crystallisation, *e.g.*, zinc sulphate, iron sulphate and copper sulphate.

Certain substances have a strong affinity for water and take up moisture on exposure to air. If strong sulphuric acid or quicklime be left exposed to air for a few hours, it absorbs moisture from the atmosphere and increases in weight. Such substances are said to be **hygroscopic**, and may be used to detect the presence of water vapour in air. Certain solid substances are able to take up so much moisture that they become liquid by dissolving in the water absorbed from the atmosphere. They are known as **deliquescent** substances. If a little of caustic soda be left exposed to air for an hour or so, it will absorb sufficient water to form its aqueous solution. Caustic potash and calcium chloride are important examples of deliquescent substances. These substances are used for drying gases on which they have no chemical action.

The action of metals on water varies with the nature of the metal. Certain metals like sodium and potassium decompose it at the ordinary temperature; others like iron, zinc and magnesium do so only when heated; while gold and silver do not decompose it at all.

Experiment 45.—Throw a small piece of sodium on water in a porcelain dish. The metal floats about the surface of the water and a hissing sound is produced. As the action goes on, sodium gradually diminishes in size, and finally disappears. Touch the liquid in the dish, it feels soapy, showing the formation of an alkali. (See Chapter VIII, page 68).

Repeat the experiment with potassium. The action is so vigorous that the liberated hydrogen gas catches fire. The violet colour of the flame is due to the potassium

vapour evolved at the same time.

When steam is passed over heated magnesium, zinc, or iron, it is decomposed into its elements. The oxygen combines with the metal, forming the corresponding oxide, while hydrogen is evolved in the free state (see expt. 48, page 58).

QUESTIONS.

1. A colourless liquid is given to you, what tests would you make to find that the given liquid is pure water?

2. Name some places in your district where water appears to have a greenish-blue colour.

3. How would you show that water is a chemical compound and not a mixture of hydrogen and oxygen?

4. Name the chief impurities present in (*i*) rain-water, (*ii*) spring-water, and (*iii*) river-water. What are the sources of these impurities?

5. How does hard water differ from soft water? What is the hardness of water due to? How can permanent hard water be softened?

6. What are mineral waters? Name some important mineral springs in your province.

7. In what respects do rain-water, river-water, and sea-water differ from one another? How would you prepare fresh water from a dirty sample of sea-water? (P. U. Matric., 1923.)

8. What experiments would you make to distinguish *hard* water from *soft* water? What steps would you take to remove the *temporary* hardness of water? (P. U. Matric., 1922.)

9. How would you explain the formation of 'fur' in kettles and boilers?

10. How can calcium carbonate be made to dissolve in water, and how can it be recovered from the solution so obtained?

CHAPTER VII.

HYDROGEN AND COMPOSITION OF WATER.

Hydrogen.

39. **History.**—It has been known for a very long time that an inflammable gas is produced by the action of sulphuric acid on iron. Cavendish in 1766 showed the gas to be an element and named it *inflammable air*. It is now called *hydrogen* on account of its presence in water. The word *hydrogen* means water-producer.

40. **Occurrence.**—Hydrogen occurs in the free state in a very small quantity in the gases given off from volcanoes, and in large quantities in the gaseous envelope around the sun. In combination with other elements, it occurs very abundantly on the earth. It is chiefly found in combination with oxygen in water of which it forms one-ninth part by weight. It is also present in kerosene oil, and in animal and vegetable substances. It is an essential constituent of all acids.

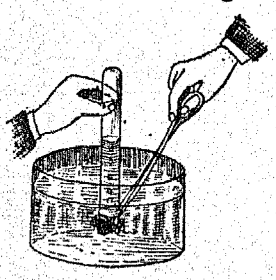


Fig. 21.—Preparation of hydrogen by sodium.

41. **Preparation.**—Hydrogen is mostly obtained from two important sources—water and acids.

I. **From water.**—(1). Sodium and potassium decompose water at the ordinary temperature and liberate hydrogen.

Experiment 46.—Take a piece of sodium in a fine wire gauze and fold the gauze. Press it under water with a pair of tongs, and invert a test-tube full of water over the gauze (Fig. 21). Bubbles of gas rise and collect in the tube. When the action stops, remove the tube and apply a burning match to its mouth. The gas burns with a pale-blue flame.

The action of sodium on water is very vigorous and cannot be controlled easily. In order to moderate the action, the metal is mixed with mercury to form *sodium amalgam*. Mercury has no action on water, and so the action of sodium amalgam on water is much slower than that of sodium alone.

Experiment 47.—Take a small quantity of mercury in a mortar. Add a small piece of sodium and rub the two together with a pestle; a flash of light is produced on rubbing. Place another piece of sodium and rub again. Repeat this till the mixture becomes solid; this is *sodium amalgam*. Put the amalgam in a small porcelain crucible. Place the crucible in a big beaker and cover it with a small inverted funnel, the end of which is 2 or 3 inches below the top of the beaker. Fill the beaker with water and invert a test-tube full of water over the funnel (Fig. 22). Bubbles of hydrogen slowly rise and collect in the test-tube. When the tube is full, remove it from the water and apply a lighted match. The gas burns with a pale blue flame.

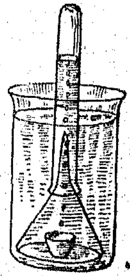


Fig. 22.—Preparation of hydrogen by sodium amalgam.

(2). Metals like magnesium, zinc, or iron can decompose water only when steam is passed over

the heated metal. The metal takes up oxygen and forms the corresponding oxide, while hydrogen is given off—

Iron + steam = iron oxide + hydrogen.

Experiment 48.—Take an iron tube about nine inches long, and connect its one end to a flask containing water, and fit the other with a delivery tube passing under a bee-hive shelf in a trough of water (Fig. 23). Place some iron filings in the tube and heat strongly. Boil the water in the flask and allow the steam to pass over heated iron filings. The iron takes up the oxygen of the steam, forming iron oxide, and leaves the hydrogen which is collected in a test-tube placed over the bee-hive shelf. Remove the tube when full of the gas and apply a burning taper. The gas burns with a pale blue flame

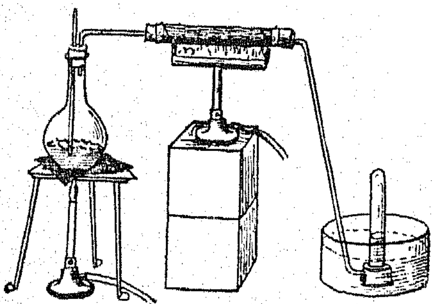


Fig. 23.—Preparation of hydrogen from steam.

Repeat the above experiment with magnesium powder, replacing the iron tube by a hard glass tube. The heated metal catches fire in the steam, forming magnesium oxide and giving off hydrogen—

Magnesium + steam = magnesium oxide + hydrogen.

(3). Pure hydrogen may be obtained by the decomposition of water, acidified with a few drops of sulphuric acid, by an electric current. The apparatus used is known as *voltameter*, and is described in experiment 53.

This method is employed for preparing hydrogen gas on a commercial scale where electric

current is cheap. In fact, most of the commercial hydrogen is now manufactured by the electrolysis of water.

II. **From acids.**—Hydrogen is usually prepared in the laboratory by the action of dilute sulphuric (or hydrochloric acid) on zinc, when zinc sulphate (or zinc chloride) is formed and hydrogen gas evolved—

Zinc+sulphuric acid
= zinc sulphate+hydrogen.

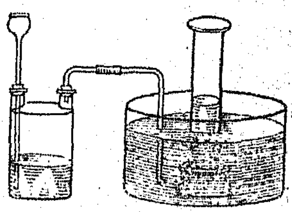


Fig. 24.—Preparation of hydrogen.

Experiment 49.—Take a Woulfe's bottle and place some granulated zinc in it. Fit one neck of the bottle with a thistle funnel and the other with a delivery tube (Fig. 24). Pour some water in the bottle through the funnel such that the zinc pieces are completely immersed in it and the lower end of the funnel dips under water. Now add dilute sulphuric acid to the bottle. A vigorous action takes place and bubbles of gas escape from the end of the delivery tube. Allow the bubbles to escape for a minute or two so as to expel the air of the apparatus. Collect the gas in a test-tube over water and test it with a lighted match. If the gas burns with an explosive noise, the hydrogen is mixed with air. Collect the gas again and keep on testing it till it burns quietly. Now collect the gas in several jars and place them inverted on the table to test the properties of the gas (see below).

NOTE.—*The gas forms an explosive mixture with air, and hence the apparatus used must be air-tight and no flame should be brought near it.*

42. **Properties.** *Physical.*—Hydrogen is a colourless, odourless and tasteless gas. The gas

prepared in expt. 49 has a faint disagreeable smell on account of the presence of impurities in commercial zinc and sulphuric acid. It is considerably lighter than air. In fact, it is the lightest substance known.

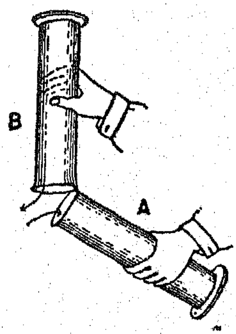


Fig. 25.—Transfer of **hydrogen.**

Experiment 50.—(*a*). Take a jar *A* of hydrogen and hold it under an inverted jar *B* of air, as shown in Fig. 25. Remove the upper jar, and holding it in the inverted position apply a burning taper to it. The gas in the jar burns with an explosive noise, showing that hydrogen is lighter than air and has passed into the upper jar *B*. Now introduce a burning taper in *A*. The gas does not burn, showing that there is no hydrogen left in the jar.

(*b*). Fit up the apparatus for generating hydrogen, and dip the end of the delivery tube in a soap solution. Soap bubbles filled with the gas are formed and rise up in the air.

(*c*). Fill a toy balloon with the gas and close its mouth with a light thread. Hold the other end of the thread in your hand and leave the balloon in air; it rises up and floats in air.

Chemical.—Hydrogen burns itself but does not allow other substances to burn in it. When the gas burns in air, it combines with oxygen to form water.

Experiment 51.—Introduce a burning candle in a jar of hydrogen held with its mouth downward. The gas burns with a pale-blue flame at the mouth of the jar, but the candle is extinguished. Take out the candle and see that it is lighted up again by the burning hydrogen.

Experiment 52.—Fit up the apparatus for the preparation of hydrogen gas (see experiment 49), using a jet tube in place of the delivery tube (Fig. 26). Invert a test-tube over it and let it stand for a minute or two. Take the test-tube aside and keeping it in an inverted position apply a burning match. If the gas burns with a slight explosion, light up the gas escaping from the jet tube. The gas burns with a pale-blue flame—*Philosopher's candle.*

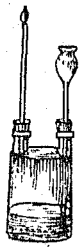


Fig. 26.—Burning of hydrogen in air.

Now invert an empty jar over the lighted jet (Fig. 27). Drops of water soon appear on the inside of the jar.

The formation of water by the burning of hydrogen in air may also be shown by holding a flask containing cold water over the lighted jet of the gas. Drops of water appear on the outside of the flask and may be collected.

When hydrogen is passed over heated copper oxide, it takes up oxygen and forms water, while metallic copper is left behind—

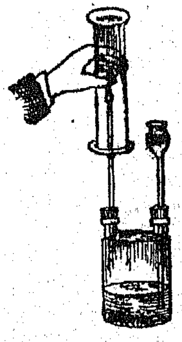


Fig. 27.—Formation of water by burning hydrogen.

Copper oxide + hydrogen = copper + water.

This process of the removal of oxygen from a compound by hydrogen (or other elements, *e.g.*, carbon) is known as **reduction**, and the compound which loses oxygen is said to be *reduced*. Thus,

copper oxide in the above example is reduced to copper.

43. **Uses of hydrogen.**—Hydrogen was formerly used in filling balloons, but the cheaper coal-gas is now used in its place. It is also used as a source of heat and light in the oxy-hydrogen flame for producing lime-light, and melting platinum in places where electric current is not available. Its chief commercial use is in the 'hydrogenation of oils,' *i.e.* the conversion of liquid oils into solid fats some of which, such as vegetable ghee (*banaspati ghee*) are so largely used now-a-days in cooking.

Composition of Water.

44. Water is composed of oxygen and hydrogen in the ratio of one to two by volume, and eight to one by weight. The composition of water by volume can be determined either by **analysis,** *i.e.* decomposing it into its component parts, or by **synthesis,** *i.e.* preparing it by the combination of its elementary constituents.

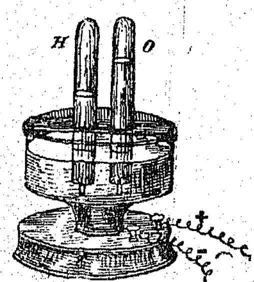


Fig. 28.—Electrolysis of water.

I. **Composition by volume.**—The ratio by volume in which hydrogen and oxygen are present in water is easily determined by decomposing it by an electric current and collecting the gases separately.

Experiment 53.—Take a glass vessel fixed on a wooden base and having two platinum wires fused at its bottom. The ends of the wires in the vessel terminate in small strips of platinum, known as *electrodes;* while

the ends outside the vessel are connected to binding-screws at the base. This apparatus is known as the **voltameter.** Fill nearly three-fourths of the voltameter with water, and add a few drops of strong sulphuric acid to it. Invert two test-tubes of the same size and shape and full of water over the electrodes. Now connect the binding-screws with the poles of a battery consisting of two of these cells (Fig. 28). Bubbles of gas immediately begin to rise from each electrode and collect in the tubes. After the current has passed for some time, the volume of the gas in one tube will be found to be double of that in the other. When one of the tubes is full, disconnect the battery. Remove the tube full of the gas and test it with a burning match. The gas burns with a pale-blue flame, showing that it is hydrogen. Now remove the other tube which is only half-full, and introduce a glowing splinter in it. The splinter bursts into flame, showing that it is oxygen.

Pure water is a bad conductor of a current of electricity, and a little sulphuric acid is added to increase its conducting power.

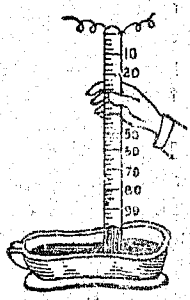


Fig. 29.—The eudiometer tube.

The electrode, which is connected with the negative pole of the battery and on which hydrogen is collected, is known as the **negative electrode** or **cathode**; and the other, which is connected with the positive pole of the battery and on which oxygen is collected, is called the **positive electrode** or **anode.** Electric current enters the acidified water in the voltameter through the anode and leaves it through the cathode.

The composition of water by volume can be more accurately determined by taking oxygen and

hydrogen in the ratio of one to two in a **eudiometer tube** standing over mercury in a small trough. The eudiometer tube is a graduated glass tube open at one end and having two short platinum wires fused near the closed end (Fig. 29). On passing electric sparks through the mixture, the gases combine and form water. When the tube cools, mercury rises up and fills the tube almost completely.

Two volumes of hydrogen combine with one volume of oxygen to form water which, at the ordinary temperature, occupies practically no volume as compared with the volume of the mixture. If, however, the eudiometer tube be heated to 100°C., the liquid water will be converted into steam which occupies two-thirds of the original volume of the mixture of hydrogen and oxygen.

Example.—30 c.c. of hydrogen are mixed with 20 c.c. of oxygen in a eudiometer tube and the mixture subjected to electric sparks. Find the nature and volume of the gases left in the tube at (*i*) 100°C., (*ii*) room temperature.

We know that 2 vols. of hydrogen combine with 1 vol. of oxygen to give 2 vols. of water vapour at or above 100°C. This vapour condenses to water below 100°C. and occupies a volume which is so small as to be negligible. Now—

Hydrogen +	oxygen =	water vapour =	water
2 vols.	1 vol.	2 vols.	0 vol.

$\therefore$ 30 c.c. of hydrogen would combine with 15 c.c. of oxygen to form 30 c.c. of water-vapour at 100°C.

Hence the gases left in the tube at 100°C. are (*i*) $20 - 15 = 5$ c.c. of oxygen, and (*ii*) 30 c.c. of water vapour.

If, however, the temperature is below 100°C., 30 c.c.

of water vapour will condense to water and occupy zero volume.

Hence the gas left in the tube at room temperature will be 20 — 15 = 5 c.c. of oxygen.

II. **Composition by weight.**—The composition of water by weight can be found by passing hydrogen, dried by passing through a bottle S containing strong sulphuric acid, over copper oxide contained in a hard glass tube B (Fig. 30). The

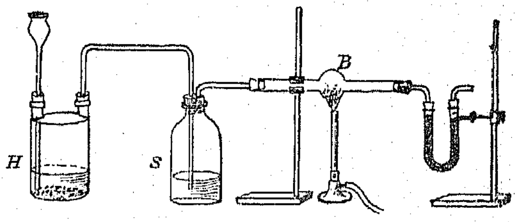


Fig. 30.—Composition of water by weight.

hydrogen combines with the oxygen of the copper oxide to form water, which is collected in a U-tube containing calcium chloride. The increase in the weight of the U-tube gives the weight of the water formed. The hard glass tube is weighed before and after the experiment, and the loss in weight is the weight of the oxygen taken up by the hydrogen to form water.

Suppose—

Weight of hard glass tube and copper oxide
(before the experiment) .. .. = 30·0 gms.
Weight of hard glass tube and
copper oxide (after the experiment) = 28·4 gms.
∴ Weight of oxygen lost.. .. = 1·6 gms.
Weight of U-tube and calcium chloride

(before the experiment) .. .. = 40·0 gms.
Weight of U-tube and calcium chloride
(after the experiment) .. .. = 41·8 gms.
∴ Weight of water formed .. .. = 1·8 gms.
Hence weight of hydrogen = 1·8 — 1·6 gms.
= ·2 gm.

Therefore the ratio between the weights of oxygen and hydrogen forming water is 16 : 2, or 8 : 1.

QUESTIONS.

1. Describe any two methods by which hydrogen may be obtained from water. Sketch the apparatus used in each case.

2. How is hydrogen prepared in the laboratory? What precautions are necessary for collecting the gas?

3. Describe an experiment to show that water is formed when hydrogen burns in air.

4. Give a short account of the preparation and properties of hydrogen gas.

5. Describe experiments to prove that hydrogen is lighter than air.

6. When zinc is placed in hydrochloric acid, the zinc gradually disappears. Sodium chloride also disappears when put into water. Explain the difference between the changes that take place in the two cases.

7. An electric spark is passed through a eudiometer tube filled with equal volumes of hydrogen and oxygen. What substances will be found in the eudiometer after the explosion? What is the volume of each substance after cooling? [P. U. Matric., 1918.]

8. 120 c.c. of hydrogen are mixed with 40 c.c. of oxygen in a eudiometer tube and the mixture subjected to an electric discharge. What volume of gas is left behind, and what is the nature of this gas?

9. How is the composition of water determined by means of an electric current? Draw a diagram of the apparatus you would employ.

10. Which metals decompose water at (*i*) ordinary temperature, and (*ii*) red heat?

11. You are provided with four jars containing air, oxygen, nitrogen and hydrogen. What experiments would you make to distinguish these gases from one another? [P. U. Matric., 1924.]

CHAPTER VIII.

ACIDS, BASES AND SALTS.

45. It has already been mentioned on page 44 that there are 3 kinds of oxides: *acidic oxides*, which combine with water to form acids; *basic oxides*, which form bases, and *neutral oxides*, which have no noticeable action on litmus. Acids and bases combine to form compounds, known as *salts*. These three classes of compounds, namely acids, bases and salts, possess certain characteristic properties by which they can be easily distinguished from one another.

46. **Acids.**—The most important acids, which are commonly used both in commerce and the laboratory, are:—

No.	*Name of the acid.*	*Composition.*
1.	Sulphuric acid.	Sulphur, hydrogen and oxygen.
2.	Nitric acid.	Nitrogen, hydrogen and oxygen.
3.	Hydrochloric acid.	Hydrogen and chlorine.

The individual properties and uses of each of these acids will be discussed in due course. In this chapter we shall take up only those properties which are commo to all the acids.

The characteristic properties of acids are:—

1. *Hydrogen is an essential constituent of all acids.*

2. *Acids have a sour taste.*

Experiment 54.—Take a small beaker and nearly fill it with water. Add a drop of an acid, say sulphuric acid, to the water and stir the liquid. Taste this liquid; it will be found to have a sharp sour taste.

3. *Acids turn blue litmus red.*

Experiment 55.—Take some water in a beaker and add to it about one-fourth its volume of an acid. Now add a few drops of this dilute acid to some blue litmus solution contained in a test-tube. The blue colour at once changes to red.

4. *Acids react with certain metals, giving hydrogen.* For example, dilute sulphuric acid readily dissolves zinc or iron with the evolution of hydrogen (page 58).

5. *Acids decompose carbonates, giving off carbon dioxide.*

Experiment 56.—Take a little sodium carbonate (washing soda) in a test-tube and add a few drops of dilute hydrochloric acid to it. A strong effervescence takes place due to the evolution of carbon dioxide—

Sodium carbonate + hydrochloric acid
= sodium chloride + water + carbon dioxide

47. **Bases.**—A base is a substance which combines with an acid to form a salt and water. Bases are usually the oxides or hydroxides (compounds of oxides and water) of metals. Their solubility in water varies considerably. Those, which dissolve in water and impart a soapy touch and a bitter taste to it, are termed **alkalies.** The names and composition of the most important alkalies are given in a tabular form on the next page.

The characteristic properties of alkalies are:—

1. *All alkalies contain both hydrogen and oxygen. Alkalies have a soapy touch and a bitter taste.*

No.	Name of the alkali.	Composition.
1.	Caustic soda or sodium hydroxide.	Sodium, hydrogen and oxygen.
2.	Caustic potash or potassium hydroxide.	Potassium, hydrogen and oxygen.
3.	Ammonium hydroxide.	Nitrogen, hydrogen and oxygen.
	Calcium hydroxide.	Calcium, hydrogen and oxygen.

Experiment 57.—Dissolve some caustic soda or potash in water, and touch the solution with your fingers. On rubbing the fingers, a soapy feeling is produced.

Dilute a little of the above solution with a sufficient quantity of water. Taste the solution; it has a bitter taste.

2. *Alkalies turn red litmus blue.*

Experiment 58.—Take some blue litmus solution in a test-tube and turn it red by the addition of a drop of an acid. Now add caustic soda solution to it; the blue colour reappears.

3. *Alkalies turn yellow turmeric brown.*

Experiment 59.—Take a little powdered turmeric (*haldi*) and dissolve it in hot water. Add a few drops of caustic soda solution to the turmeric solution; the yellow colour changes to brown.

Note.—Instead of litmus and turmeric solutions, litmus and turmeric papers are generally used in the laboratory. They are prepared by dipping a paper in the corresponding solution and drying the paper.

4. *Alkalies neutralise acids with the formation of salts.*

When an alkali is mixed with an acid in proper proportions, the characteristic properties of both the alkali and the acid are destroyed, and the two are said to *neutralise* each other. The pro-

ducts formed are *neutral*, and the process is known as **neutralisation.**

Experiment 60.—Take a solution of caustic soda and add a little red litmus solution to it; the colour at once changes to blue. Now add a small quantity of an acid, say hydrochloric acid, drop by drop; no change in colour is observed in the beginning. But as more of the acid is added, a point is reached when the colour is neither blue nor red. The solution is then *neutral.* At this point the addition of even a drop of an acid will change the colour of the solution to red, while the addition of an alkali will turn it blue.

Evaporate the neutral solution to dryness. A solid residue is left behind, and it has none of the properties of the acid or the base. Caustic soda and hydrochloric acid combine together and form a neutral substance, sodium chloride (common salt)—

Caustic soda + hydrochloric acid
= sodium chloride + water.

48. **Salts.**—A salt is a substance produced by the action of an acid on a base. Thus, common salt is obtained by neutralizing caustic soda with hydrochloric acid. The salts obtained from sulphuric acid are known as **sulphates,** those from nitric acid are **nitrates,** while those from hydrochloric acid are **chlorides.**

Salts may be prepared in several ways:—

1. *By the action of an acid on a metal.*—Sulphuric acid reacts with zinc, forming zinc sulphate—

Zinc + sulphuric acid = zinc sulphate + hydrogen.

Experiment 61.—Take a few pieces of granulated zinc and add dilute sulphuric acid to it. A vigorous action takes place and zinc goes into solution, giving off hydrogen gas. When the whole of the zinc has dissolved, pour the solution in an evaporating dish. Concentrate the solution by heating, and allow it to stand.

Crystals of zinc sulphate are obtained.

2. By the action of an acid on an oxide, hydroxide, or carbonate of a metal.—Lead oxide (litharge), when treated with nitric acid, forms lead nitrate—

Lead oxide + nitric acid = lead nitrate + water.

Experiment 62.—Take some litharge (*murda sang*) in a porcelain dish, add nitric acid and warm. As the oxide dissolves in the acid, its colour disappears. Evaporate the solution to dryness; lead nitrate is left behind in the form of a white solid.

The formation of a salt by the action of an acid on a hydroxide has already been described in expt. 60.

Potassium nitrate is formed when nitric acid reacts with potassium carbonate—

Potassium carbonate + nitric acid
= potassium nitrate + water + carbon dioxide.

Experiment 63.—Take potassium carbonate in a beaker and add enough of nitric acid to it. A strong effervescence takes place due to the evolution of carbon dioxide. When the whole of the carbonate has dissolved, transfer the solution to an evaporating dish. Concentrate the solution and allow it to stand. Crystals of potassium nitrate (nitre) are formed.

3. By the action of an acid on a salt of another acid.—Sodium sulphate is formed when sodium chloride is mixed with concentrated sulphuric acid and heated—

Sodium chloride + sulphuric acid
= sodium sulphate + hydrochloric acid.

Experiment 64.—Take some sodium chloride in a dish and add sufficient sulphuric acid to it. Heat the dish and stir the mixture. White fumes with a strong pungent smell are evolved. When the whole of the salt has dissolved, concentrate the solution and let it cool.

White crystals of sodium sulphate will be formed.

Unlike acids and alkalies, salts are not characterised by the presence of any particular element or group of elements. They have no particular taste as a class, and are generally without any action on litmus.

QUESTIONS.

1. What do you understand by the terms: *acid, base, alkali* and *salt?* How will you distinguish one from the other? Give examples of each.

2. What are the characteristic properties of acids? How do they differ from those of alkalies?

3. What do you understand by the terms: *alkali* and *alkaline reaction?* Name the three important alkalies.

4. What happens when sodium is brought into contact with water? How does the liquid that remains differ from water?

5. What is the difference between an oxide and a hydroxide? What is formed when an acid acts on each of them? Give examples.

CHAPTER IX.

CARBON.

49. Occurrence.—Carbon is very widely distributed in nature, and occurs both in the free and combined states. In the free state, it is found as diamond and graphite—its crystalline forms, and as charcoal—its non-crystalline or amorphous form. In combination with oxygen, it occurs as carbon dioxide in air (page 33) and in natural waters (page 51). When combined with a metal and oxygen, it occurs abundantly as carbonates, such as limestone, marble and chalk. In combination with hydrogen, carbon is found in kerosene oil, paraffin wax, marsh-gas and coal-gas. It is an essential constituent of all animal and vegetable substances. In fact, carbon is present in a larger number of compounds than any other element.

50. Forms of carbon.—Carbon occurs in nature in various forms:—

Fig. 31.—The *Koh-i-noor* diamond (natural size).

1. **Diamond.**—This is the purest form of carbon and is found in India, Brazil, South Africa and Australia. The diamonds found in the Golconda mines (Madras), especially the famous *Koh-i-noor* (Fig. 31), have long been prized for their purity and brilliancy. Diamonds may be colourless, or of a yellow, pink

blue, green, or black colour due to the presence of traces of impurities. They are artificially prepared by dissolving pure charcoal in molten iron at a very high temperature and cooling the molten mass suddenly.

Diamond, when pure, is a colourless, transparent and crystalline substance. It is the hardest substance known, and can cut all other substances. It is a bad conductor of heat and electricity. It is the densest form of carbon, having a specific gravity of 3·5. Its refractive index is higher than that of any other transparent solid, and hence its value as a gem. When strongly heated in oxygen, it burns without melting and forms nothing but carbon dioxide, showing that diamond is a pure form of carbon.

Black diamonds, commonly known as *carbonado*, are of no value as gems, and are used in cutting glass and in grinding and polishing hard substances.

2. **Graphite** (*black-lead or plumbago*).—Graphite occurs in large quantities in California, India and Ceylon. It is a soft, greyish-black substance with a metallic lustre and a soapy touch. Its specific gravity is about 2·3. Unlike diamond, it is a good conductor of heat and electricity, and is therefore used in forming electrodes. It is so soft that it marks the paper, and when mixed with clay it is used in making pencils. These pencils are commonly known as 'lead pencils,' for they were wrongly supposed to contain lead.

Graphite is very infusible, and is consequently used along with fire-clay in making 'plumbago' crucibles. It is also employed as a coating for

iron to prevent rusting, and as a lubricant for machinery where oil cannot be used. When strongly heated, graphite burns in oxygen, forming carbon dioxide and leaving a small residue.

Experiment 65.—Mix some powdered graphite with twice its weight of potassium chlorate, and apply a flame to the mixture. The graphite takes up oxygen from the potassium chlorate and burns with a brilliant white light.

51. **Amorphous carbon.**—The most familiar forms of amorphous carbon are: wood charcoal, animal charcoal, lampblack, coke and coal.

1. **Wood charcoal.**—This is prepared by strongly heating wood out of contact with air. In the Punjab, the method generally employed consists in heaping up a pile of wood in such a way as to leave a central space for a chimney and a few air holes at the bottom. The pile is then covered with earth and ignited at the bottom. The wood is allowed to burn slowly, the necessary air being supplied through the air-holes. When the fire has spread throughout the pile, the air-holes are closed and the fire is thus extinguished. The wood continues to smoulder and is finally converted into charcoal. The process is wasteful, as a part of the wood is burnt away and all the volatile products escape.

A more economical method is to heat strongly small logs of wood in large iron retorts fitted with the condensing arrangement. The volatile products distil over, and charcoal is left in the retorts. This process is known as the **destructive distillation** of wood.

Wood charcoal is a black, porous substance having a specific gravity of about 1·5. It floats on water on account of the air enclosed within its

pores. If this air be removed by immersion in boiling water, the charcoal sinks down.

Experiment 66.—Tie a piece of lead or some other metal to a charcoal piece. Place it in a beaker of water such that the charcoal dips under water. Boil the water for some time. Let it cool and cut the thread connecting the lead and charcoal. The charcoal no longer floats on water. The boiling water has expelled all the air enclosed within the pores of the charcoal which consequently sinks down.

Charcoal is a bad conductor of heat and electricity. It is used as a fuel, and burns in air without smoke, leaving behind a small quantity of ash.

It has a remarkable property of absorbing gases, and is therefore employed in deodorising and purifying water (page 12) and air.

Absorption of gases by charcoal can be easily shown by the following experiment.

Experiment 67.—Collect some dry ammonia gas in a wide glass test-tube and invert it over mercury in a dish (Fig. 32). Hold a piece of charcoal in a pair of tongs and heat it so as to expel the enclosed air. Now introduce the charcoal in the tube. The gas is readily absorbed by the charcoal and the mercury rises up in the tube.

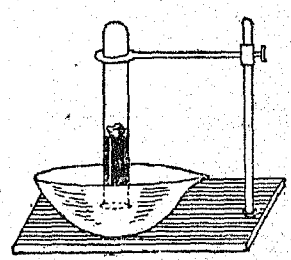


Fig. 32.—Absorption of ammonia by charcoal.

Charcoal resists the action of moisture, and hence the lower ends of fence posts, which are to be buried underground, are sometimes charred at the surface.

A table comparing the physical properties of diamond, graphite and charcoal is given below:—

No.	Property.	Diamond.	Graphite.	Charcoal.
1.	Colour ...	When pure, colourless and transparent solid.	Greyish-black and opaque solid.	Black solid.
2.	Purity ...	Purest form of carbon.	Less pure ...	Least pure.
3.	Hardness ...	Hardest substance known.	Soft and soapy to the touch.	Less soft than graphite.
4.	Specific gravity.	3·5	2·3	1·5
5.	Conductivity ...	Bad conductor of heat and electricity.	Good conductor.	Bad conductor
6.	Absorption of gases.	Does not absorb any gas.	Cannot absorb any appreciable quantity of any gas.	Can absorb large quantities of gases on account of its porous nature.

2. **Animal charcoal** (*bone-black*).—Like wood charcoal, this form of carbon is obtained by the destructive distillation of bones. It contains only about 19 per cent. of carbon, the remaining being mainly bone-ash (calcium phosphate). It is a much better absorbent of gases and colouring matter than wood charcoal, and is largely used for refining brown sugar.

Experiment 68.—Dissolve some brown sugar (*gur*) in warm water; a brown solution is formed. Add a little of powdered animal charcoal and boil the solution. Allow the mixture to stand and filter. The brown colour

is removed by the charcoal and a colourless liquid is obtained.

Repeat the experiment with litmus and indigo solutions, and note the disappearance of colour in each case.

3. **Lampblack.**—This is the purest form of amorphous carbon and is obtained as **soot** when substances rich in carbon burn in an insufficient supply of air. On a large scale it is prepared by burning impure turpentine, kerosene oil, or paraffin wax in a closed space, and allowing the soot to pass into chambers where it gets deposited on coarse blankets. It is used for making black paints and Indian ink. In villages, where rape-seed oil (*karwa tel*) is still burnt in a small earthen vessel (*diwa*), a good deal of soot (*kajal*) is obtained and used in making cheap ink.

4. **Coal.**—Coal is an impure form of carbon, and exists in enormous quantities in nature. Most of the coal in the Punjab is obtained from coal mines of Dandot (Punjab) and Jharia (Bengal). It is produced in nature as a result of slow decomposition of vegetable matter in the absence of air and under deposits of sand and mud. This process of decay is supposed to have been going on for long ages and under enormous pressure. The impressions of leaves and flowers sometimes observed on the surface of coal confirm this view. Coal is extensively used as a fuel in engines, kilns, steam-boilers and for various other purposes. It is difficult to ignite, but when once ignited it continues burning, giving out intense heat without smoke. When distilled in closed retorts, it yields coal-gas which is so largely employed for burning and lighting purposes (page 80).

5. **Coke.**—This is left behind as a hard substance in the retorts when coal is distilled out of contact with air. It is largely used as a household fuel, and as a reducing agent in obtaining metals from their ores.

All forms of carbon are chemically the same, because they all burn in oxygen and give carbon dioxide.

52. **Chemical properties of carbon.**—Carbon combines directly with many elements at high temperatures, though it is extremely inactive at the ordinary temperature. Its compounds with metals are known as **carbides,** of which calcium carbide, so largely used in the preparation of acetylene gas, is an important example. When calcium carbide is treated with water, calcium hydroxide is formed and acetylene gas is evolved—

Calcium carbide+water
=calcium hydroxide+acetylene.

When sulphur vapour is passed over red-hot charcoal, carbon disulphide is formed—

Carbon+sulphur = carbon disulphide.

Carbon has a strong affinity for oxygen and is therefore used as a reducing agent in the extraction of several metals, *e.g.* zinc, iron and tin from their oxide ores.

53. **Coal-gas.**—This gas is obtained by distilling coal in closed iron retorts, when coal-tar and ammoniacal liquor distil over and are condensed as liquids. Coal-gas passes over and is collected in gas-holders, while coke is left behind as a hard solid substance in the retorts. Coal-tar is a dark-brown oily liquid. It forms an important source of several substances like benzene, carbolic acid

and aniline dyes. Ammoniacal liquor is the chief source of ammonia and ammonium compounds. When the liquor is mixed with excess of lime and heated, ammonia gas is given off. On passing the gas in different acids, ammonium salts are obtained.

Coal-gas may be prepared in a small quantity by heating coal in a hard glass test-tube.

Experiment 69.—Take a hard glass test-tube and fit it with a cork and a delivery tube, as shown in Fig. 33. Spread a thin layer of dry powdered coal in the tube and heat at first gently and then strongly. Smoke and brown fumes are seen in the delivery tube. A thick, dark-brown, oily liquid, *coal-tar,* and an upper watery layer, *ammoniacal liquor,* are collected in the bottle due to condensation of the vapour evolved from the coal.

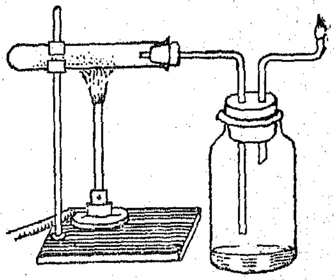


Fig. 33.—Preparation of coal-gas.

Apply a burning match to the gas coming out of the jet tube; it burns. This is coal-gas. The hard black substance left behind in the test-tube is coke.

54. **Oil-gas.**—The gas generally used in the science laboratory of certain high schools and in almost every college is not coal-gas but oil-gas obtained by the *cracking* (*i.e.* breaking up at a high temperature) of kerosene oil in closed iron retorts. In some schools where a small quantity of the gas is required for heating or burning purposes, petrol gas is used. It is formed by the *cracking* of petrol. The plant for petrol gas is smaller, cheaper and more convenient to work than the one employed for preparing the oil-gas.

QUESTIONS.

1. Give a brief description of the different forms of carbon.
2. What is amorphous carbon? Give a brief account of the various forms of amorphous carbon.
3. How is charcoal prepared from wood? Describe the properties and uses of charcoal.
4. How is coal formed in nature and what are its uses?
5. What is meant by *destructive distillation?* What are the products obtained by the destructive distillation of coal?
6. What is animal charcoal? How is it prepared, and for what purposes is it used?
7. How would you show that diamond, graphite and charcoal consist essentially of the same element?

CHAPTER X.

OXIDES OF CARBON.

Carbon Dioxide.

55. **History.**—In the seventeenth century it was observed that the gas given off by burning wood or by the action of acids on chalk differed from air in being a non-supporter of combustion. Later on, the gas was shown to be a constituent of chalk and termed *fixed air*. A few years after, Lavoisier, a French chemist, prepared it from carbon and oxygen and named it carbon dioxide.

56. **Occurrence.**—As we have already seen, carbon dioxide is present in the atmosphere and in natural waters in varying proportions. Its presence in air is due to the combustion of fuel, respiration of animals, and the decay of vegetable matter. In the combined form, it exists in large quantities in limestone, chalk, marble and other carbonates.

57. **Preparation.**—(1). Carbon dioxide is formed when carbon burns in air or oxygen (see experiments 36 and 37).

(2). It is obtained in the pure state by heating chalk (calcium carbonate), when carbon dioxide is evolved and calcium oxide left behind—

Calcium carbonate=calcium oxide+carbon dioxide.

This method is sometimes used in preparing the gas on a large scale. Limestone is burnt in special kilns, and the gases coming out consist mainly of carbon dioxide.

(3). In the laboratory, the gas is readily prepared by the action of dilute hydrochloric acid on

chalk or marble, when calcium chloride is formed and carbon dioxide given off—

Calcium carbonate+hydrochloric acid
=calcium chloride+water+carbon dioxide

Experiment 70.—Put some pieces of marble in a Woulfe's bottle fitted with a thistle-funnel and a delivery tube bent twice at right angles (Fig. 34). Pass the vertical limb of the delivery tube, which goes into the gas jar, through a cardboard disc which serves as a cover for the jar. Pour sufficient water in the bottle through the funnel so that the marble pieces are covered with water and the lower end of the funnel dips in it. Next add strong hydrochloric acid. A brisk effervescence takes place, and the carbon dioxide evolved collects in the jar by displacing the air. This method of collecting the gas is known as *downward displacement*, because the gas goes downward and displaces the air.

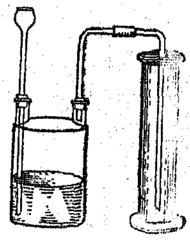


Fig. 34.—Preparation of carbon dioxide.

Bring a burning match near the mouth of the jar. If it is extinguished, the jar is full of the gas. Collect a few jars and test the properties of the gas (see below).

58. **Properties.** *Physical.*—Carbon dioxide is a colourless gas with a faint smell and an acid taste which is easily detected in 'soda water.' It is fairly soluble in water at the ordinary temperature and pressure, and cannot therefore be collected over water. It is much more soluble in water under high pressure. Soda-water, lemonade and other aerated waters are solutions of carbon dioxide in water under pressure, and the gas escapes as soon as the pressure is released on opening the bottle. Carbon dioxide is about $1\frac{1}{2}$ times as

heavy as air, and is therefore collected by downward displacement of air.

Experiment 71.—Take a jar of the gas and turn it upside down for a minute. The heavy gas falls out and air takes its place. Bring a burning candle in the jar; it is no longer extinguished.

Experiment 72.—Take a jar *A* of the gas and holding its mouth slantwise near a jar *B* of air, pour the gas in the latter as you pour a liquid from one vessel to another (Fig. 35). Remove the upper jar *A* and bring a lighted candle in it. The candle continues to burn, showing the absence of carbon dioxide in the jar. Now introduce the lighted candle in the jar *B*. It is soon extinguished, showing that carbon dioxide has been poured into it from *A*.

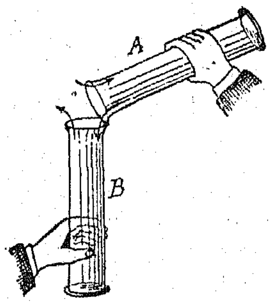


Fig. 35.—Transfer of carbon dioxide.

Experiment 73.—Blow air from your mouth by means of a glass tube in a soap solution, and take a soap bubble at the end of the tube. Allow it to drop into a jar of the gas; the bubble floats in the heavy gas.

The gas condenses to a colourless liquid under high pressure and low temperature.

Chemical.—Carbon dioxide is neither combustible nor a supporter of combustion and animal life. Animals die in the gas not only for want of oxygen, but also on account of the slightly poisonous nature of the gas.

Experiment 74.—Introduce a burning candle in a jar of the gas. The candle is at once extinguished and the gas does not burn.

Experiment 75.—Light up three or four candles and place them in a row. Now take a jar of carbon dioxide and pour the gas on the lighted candles, one after

the other. The candles are extinguished by the invisible gas as if by a wand of magic.

This property of the gas enables it to be used as a fire-extinguisher.

Experiment 76.—Take a little of kerosene oil or methylated spirit in a small dish. Ignite it by bringing a flame near the liquid. Now invert a jar of carbon dioxide over the burning liquid. The flame is extinguished.

A few metals like potassium and magnesium continue burning in the gas, taking up oxygen and leaving carbon behind—

Magnesium+carbon dioxide
= magnesium oxide+carbon.

Experiment 77.—Burn a piece of magnesium ribbon and lower it in a jar of the gas. Magnesium continues burning, and a white powder (magnesium oxide) mixed with carbon particles drops at the bottom of the jar.

A solution of carbon dioxide in water is slightly acidic to litmus and turns blue litmus solution to a wine-red colour, quite unlike the bright red colour produced by hydrochloric or sulphuric acid. The gas turns lime-water milky due to the formation of insoluble calcium carbonate—

Lime-water+carbon dioxide
= calcium carbonate+water.

Experiment 78.—Prepare some lime-water, as given in expt. 31, and pour it into a jar of carbon dioxide. Shake the jar; the lime-water turns milky.

This property of carbon dioxide is used as a test for the gas.

59. **Uses of carbon dioxide.**—The use of carbon dioxide in preparing aerated waters and extinguishing fire has already been mentioned. The gas is stored in steel cylinders under pressure for use in the manufacture of aerated waters. When required

for putting out fire, it is stored up in iron cylinders, or is generated in fire-extinguishers. A common form of a fire-extinguisher consists of a metallic vessel containing washing-soda and sulphuric acid kept apart. The two substances can be easily mixed together to obtain carbon dioxide when desired. Solid carbon dioxide is used in medicine under the name *carbonic acid snow* in curing local sores.

60. **Respiration.**—We have seen that oxygen supports the combustion of ordinary substances, such as candle, wood, etc. Oxygen is required not only for the burning of these substances, but is also essential for the maintenance of animal life. Animals die in air from which oxygen is removed. Fish living in water breathe the oxygen dissolved in it. If a fish be put in water which has been boiled to expel oxygen and cooled, it soon dies for want of free oxygen.

When we breathe in air, the oxygen of the air is carried by the blood from the lungs to the heart which sends it with the blood through the innumerable blood-vessels all over the body. This oxygen is used up in oxidising the waste-matter of the body, thereby producing heat which maintains the body at the necessary temperature. The carbon and hydrogen of the waste-matter combine with the oxygen and form carbon dioxide and water respectively, which are carried with the blood through the veins back to the lungs, where they are breathed out into the air. It is for this reason that the air breathed out contains more carbon dioxide than the air breathed in.

Experiment 79.—Take a flask fitted with two tubes.

bent at right angles (Fig. 36). Add some freshly prepared lime-water to the flask so that the lower end of the longer tube dips into the liquid. Apply the lips to the short tube and suck in air through the lime-water; the lime-water remains unchanged. Now blow into the liquid from your lungs through the other tube; the lime-water soon turns milky.

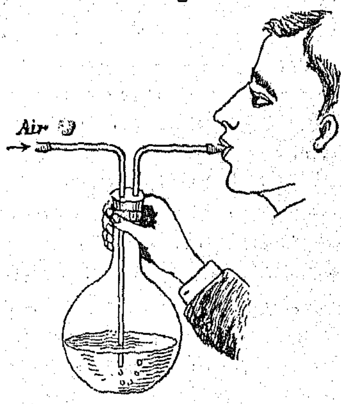


Fig. 36.—Carbon dioxide in air.

The presence of water vapour in the human breath can be easily shown by breathing on a piece of stone slab, a slate, or a mirror. Moisture is at once seen deposited on the surface.

61. **Respiration in plants.**—If the carbon dioxide continuously exhaled by animals remained in air, a time would come when it will be no longer possible for animals to live. But nature controls the proportions of the gases in the atmosphere in an excellent manner. What is poison to animals is an essential food for plants. In the presence of sunlight plants take up carbon dioxide from air. They decompose the carbon dioxide, retain the carbon to build up their tissues, and give back the oxygen to the air. This action of plants is possible only in the presence of sunlight. During the night plants revert to the same process of respiration as the animals, *i.e.* they take in oxygen and give out carbon dioxide. Thus, sunlight is an essential factor for the life and growth of plants. Since the

action of plants is the reverse of that of animals, a balance in the quantity of oxygen and carbon dioxide is maintained in air, and the proportion of the two gases in the atmosphere remains almost unaltered.

Experiment 80.—Take some green leaves of a water plant, such as hydrilla (*jala*) and place them in a large beaker full of water in which carbon dioxide has been blown (fresh 'soda-water' will do). Cover the leaves with an inverted funnel, keeping its stem below the surface of the water. Invert a test-tube full of water over the funnel (Fig. 37) and expose the beaker to sunlight for a few hours. Bubbles of gas are given off from the leaves and collect in the tube. Test the gas with a glowing splinter; it will be found to be oxygen. This oxygen comes from the carbon dioxide which is dissolved in the water and is decomposed by the leaves.

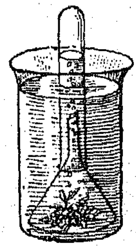


Fig. 37.—Oxygen given out by leaves.

62. **Carbonates.**—It has already been mentioned that when carbon dioxide is passed into lime-water, calcium carbonate is formed. In the same way, when carbon dioxide is passed into a solution of caustic soda, sodium carbonate is produced. Carbonates are very widely distributed in nature, and calcium carbonate is the most abundant of them all. Calcium carbonate containing calcium, carbon and oxygen occurs in several forms: limestone (*kali ka pathar*), marble (*marmar*), chalk (*khari matti*), calcite, aragonite and Iceland spar. It is insoluble in water, but it dissolves in water containing carbon dioxide due to the formation of the soluble bicarbonate—

Calcium carbonate+water+carbon dioxide
= calcium bicarbonate.

When heated strongly or treated with an acid, it decomposes to give off carbon dioxide—

Calcium carbonate=calcium oxide+carbon dioxide.

Calcium carbonate+hydrochloric acid
= calcium chloride+water+carbon dioxide.

Most of the marble in India is obtained from Jaipur (Rajputana), and is used in the construction of buildings. Pure marble is white, but the presence of impurities, especially organic matter and iron compounds, produces blue, yellow, reddish and black varieties. Limestone is found in nearly every part of India. The most important sources, however, are Jubbulpore (Bengal), Kathiawar, Vindhya and Himalaya ranges. Limestone is mainly employed in the manufacture of lime. All carbonates are decomposed by acids with the evolution of carbon dioxide.

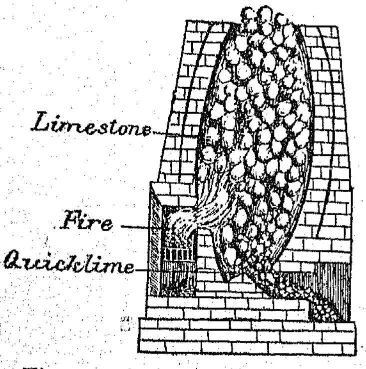


Fig. 38.—Burning of limestone in a lime-kiln.

63. **Lime.**—In the Punjab, large quantities of limestone or *kankar* are burnt in lime-kilns. A lime-kiln is an oval-shape construction of masonry work (Fig. 38). Limestone and coal or fuel-wood are placed in the kiln in alternate layers. The fuel is burnt in a free access of air entering through a large opening at

the bottom. Carbon dioxide mixed with air escapes, and the lime falls to the bottom. The process is continuous, the mixture of limestone and coal being added at suitable intervals from the top and burnt lime removed at the bottom. The lime thus obtained is known as **quicklime.**

When water is added to quicklime, a hissing sound is produced and steam is evolved. This addition of water to lime is called *slaking*, and the product is known as **slaked lime** (*bujha hua chuna*).

64. **Mortar** is a thick paste formed by mixing lime, sand (or *surkhi*) and water. On exposure to air, the water slowly evaporates and the lime absorbs carbon dioxide from the atmosphere. Mortar therefore 'sets' or hardens on exposure. The sand keeps the mass porous and prevents the formation of cracks on account of shrinkage that takes place during the 'setting' of mortar.

65. **Cement** is manufactured by burning a mixture of limestone, clay and sand in the right proportions. It has a characteristic property of setting under water and is therefore used for works under water.

Carbon Monoxide.

66. **Preparation.**—We have seen that when carbon burns in air or oxygen, carbon dioxide is produced. If, however, carbon burns in a limited supply of air or oxygen, carbon monoxide is formed at the same time. Carbon monoxide seldom occurs free in air. It may be obtained by passing carbon dioxide over red-hot charcoal, when carbon dioxide is reduced by carbon to carbon monoxide—

Carbon dioxide+carbon = carbon monoxide.

Experiment 81.—Take an iron tube and fill it losely with small pieces of charcoal. Connect one end of the tube to a carbon dioxide apparatus, and fit the other with a delivery tube passing into a trough containing a dilute solution of caustic soda (Fig. 39). Now heat the tube strongly to raise it to a red-heat. Allow a slow current of carbon dioxide to pass over the heated charcoal. Most of the carbon dioxide is converted into carbon monoxide and collected in a test-tube filled with the caustic soda solution and inverted over the bee-hive shelf. Remove the tube when full and apply a lighted candle to it; the gas burns with a blue flame.

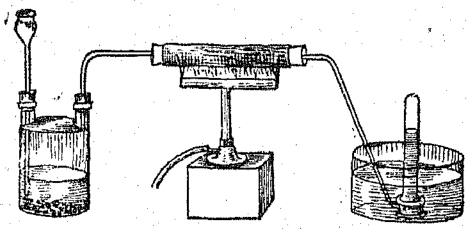


Fig. 39.—Preparation of carbon monoxide.

Caustic soda solution is used to absorb the carbon dioxide, which has not been reduced to carbon monoxide by the hot charcoal—

Caustic soda+carbon dioxide
= sodium carbonate+water.

Every one is familiar with the blue flame frequently seen over the top of a coal or charcoal-fire, when the coal or charcoal is glowing with a bright red heat. When air enters the lower parts of the grate, the charcoal burns to form carbon dioxide. This carbon dioxide passes up through the layers of red-hot charcoal, and is partially converted into carbon monoxide which burns with a blue flame at the top—

Carbon dioxide+charcoal = carbon monoxide.

67. **Properties.** *Physical.*—Carbon monoxide is a colourless, tasteless gas with a very faint smell. It is much less soluble in water than carbon dioxide. It is a very poisonous gas and produces headache even if present in small quantities in air. When inhaled in a large quantity, it causes insensibility and death. Hence it is dangerous to have a coal-fire burning in badly-ventilated rooms. In fact, many deaths have actually occurred owing to the production of the gas from coal burning in ill-ventilated or closed rooms.

Chemical.—The gas burns with a blue flame but does not support combustion. A lighted candle introduced in a jar of the gas is extinguished, but the gas burns at the mouth of the jar. Unlike carbon dioxide, it has no effect on litmus solution and does not turn lime-water milky. It has a strong tendency to take up oxygen from other substances to form carbon dioxide. For example, when passed over heated copper oxide, it takes up oxygen and reduces the oxide to metallic copper—

Carbon monoxide+copper oxide
=carbon dioxide+copper.

Carbon monoxide differs from carbon dioxide in (*i*) being combustible, (*ii*) possessing reducing property, and (*iii*) having no action on lime-water.

QUESTIONS.

1. Describe the preparation and properties of carbon dioxide. Why is the gas collected by downward displacement of air?
2. Why is carbon dioxide known as carbonic acid gas? Under what circumstances does the gas behave as an acid?
3. Describe experiments to show that carbon dioxide is (*i*) denser than air, and (*ii*) soluble in water.

4. Why do potassium and magnesium burn in carbon dioxide? Describe the action which takes place in each case.

5. Why is a strong pressure required to open a soda-water bottle? Why does the water appear to 'boil' for some time on opening the bottle? What remains after the 'boiling' has stopped?

6. Describe the action that takes place when carbon dioxide is passed through lime-water (*i*) in a small quantity, (*ii*) in excess. What happens when the liquid in the latter case is boiled?

7. You are given some lime-water and asked to blow into it. Describe the changes you observe, and name the substances produced. [P. U. Matric., 1920.]

8. Give a brief account of the changes that take place in the breathing of animals. How does breathing resemble burning?

9. Describe fully the manner in which the breathing of animals and plants affects the composition of the atmosphere.

10. How can you prove the presence of carbon dioxide in air? Describe an experiment to show the importance of carbon dioxide to plant life. [P. U. Matric., 1919.]

11. You are supplied with four jars of colourless gases: oxygen, nitrogen, hydrogen and carbon dioxide. How would you proceed to identify each gas?

12. Name some varieties of calcium carbonate and give their uses.

13. Describe the preparation of lime from limestone. How is quicklime converted into (*i*) slaked lime, and (*ii*) lime-water?

14. How would you account for the presence of blue flame seen over coal-fire?

15. How is carbon monoxide prepared? How would you obtain pure carbon monoxide from a mixture of carbon monoxide and carbon dioxide?

16. How does carbon monoxide differ from carbon dioxide? How can one be converted into the other?

CHAPTER XI.

COMBUSTION AND FLAME.

68. **Combustion.**—In Chapter V we have defined combustion as a form of oxidation attended with heat and light. The term is, however, applied in a wider sense. Any chemical action accompanied by the evolution of heat and light may be termed **combustion.** For instance, when a mixture of sulphur and iron or zinc is heated, the elements combine with the evolution of heat and light, and the process is called combustion. Similarly, the burning of hydrogen or sodium in chlorine is a case of combustion. The substance which burns is said to be **combustible**, and the gas which allows the substance to burn in it is called a **supporter of combustion.**

Every case in which heat and light are produced is not necessarily that of combustion. Thus, in an electric lamp, carbon or metallic filament gets heated and emits light when an electric current is passed through it. The glowing filament returns to its original condition as soon as the current is stopped. As no chemical change is going on in the lamp, it is not a case of combustion.

Since substances burn in air which is always present around us, the air is said to be a *supporter of combustion*, and the substances that burn in it are generally known as *combustible substances*. Thus hydrogen, carbon monoxide and coal-gas which

burn in air are generally said to be combustible gases and air the supporter of combustion. But under suitable conditions air or oxygen can be made to burn in hydrogen, carbon monoxide, or coal-gas, when the former will be combustible and the latter supporter of combustion. In fact, the terms combustible and supporter of combustion are merely relative.

69. **Flame.**—When the substances taking part in combustion are gases or vapours, **flame** is produced. Hydrogen burning in oxygen or air produces flame. The flame of a candle is in reality the flame of the vapour of wax burning in air. The wax first melts, and is then converted into vapour before it actually begins to burn. In the same way, sulphur, phosphorus, kerosene oil and rape-seed oil (*karwa tel*) are converted into vapour by the heat of burning, and the vapour then combines with the oxygen of the air, producing a flame. Carbon and iron, which do not pass into vapour by the heat of burning, burn in oxygen without any flame.

Some flames are more luminous than others. Hydrogen burns in air with an almost invisible flame, while phosphorus and magnesium burn with a dazzling light. The luminosity of a flame is generally due to the presence of solid particles heated so intensely as to emit light. For example, the oxy-hydrogen flame which is almost invisible in air, emits an intense white light, *limelight*, when a piece of quicklime is placed in it. The same principle is used in the *kitson lamp*, where the intense bright light is due to the solid mantle heated to incandescence in the flame.

The presence of solid particles is not the sole

cause of luminosity of flame. Other factors, *e.g.* density of burning gases and temperature of flame, play an important part in certain cases.

70. **The Bunsen burner.**—This is the most common form of a gas burner frequently used in the laboratory for heating purposes. It consists of a metallic tube T and side tube S near the bottom (Fig. 40). The gas (coal-gas or oil-gas) enters the burner through the side tube, and as it issues out of a small hole H, it is mixed up with air drawn in through holes at A and burns with a non-luminous flame at the top of the tube T. When a luminous flame is required, the air holes at A are closed by turning a perforated ring loosely fitting round them.

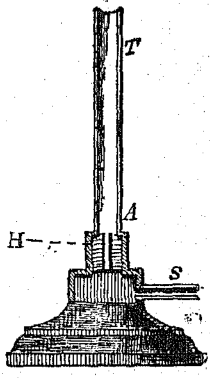


Fig. 40.—The Bunsen burner.

71. **The Primus stove.**—This stove works on the same principle as the Bunsen burner. It consists of a brass reservoir R (Fig. 41) fitted on its side with an air-pump A and at the top with a burner B having a fine hole. In order to work the stove, kerosene oil is poured into the reservoir and a little methylated spirit burnt in the cup C attached to the burner. When the spirit has burnt, air is pumped into the

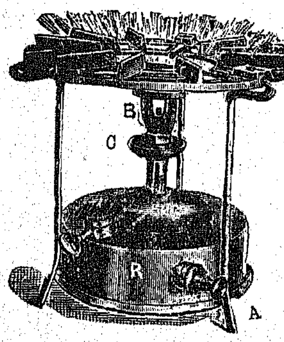


Fig. 41.—The Primus stove.

reservoir by the air-pump. This forces the oil up into the jet of the burner which, being hot, immediately vaporises the oil. The vapour of the oil, mixed with air, burns at the jet with a very hot, non-luminous flame. When the pressure of the air has diminished and flame begins to flicker, air is pumped again into the reservoir.

The Primus stove is commonly used for cooking or heating purposes.

72. **Structure of candle-flame.**—The ordinary candle-flame consists of three distinct parts (Fig. 42 A)—

1. The dark inner zone (1) surrounding the wick and consisting of unburnt vapour. The presence of unburnt vapour may be shown as follows:—

Experiment 82.—Quickly lower a piece of white paper over the flame so that the flame is slightly depressed. The paper is charred in the shape of a ring, the inside of the ring remaining unaffected.

Repeat the experiment with a wire gauze and see that a red-hot ring with a black centre is obtained.

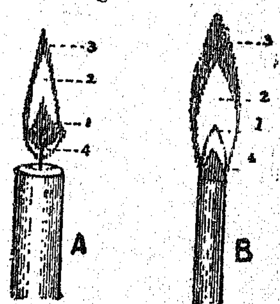


Fig. 42.—Structure of (A) candle-flame. (B) Bunsen flame.

Experiment 83.—Push the tip of a match quickly into the inner zone and hold it obliquely. The tip is unaffected, while the neighbouring part begins to burn and finally lights up the match.

Experiment 84.—Place one end of a glass tube in the inner zone as shown in Fig. 43, and apply a burning match at the other end. The unburnt vapour passes up through the tube and catches fire.

2. The luminous zone (2) sur-

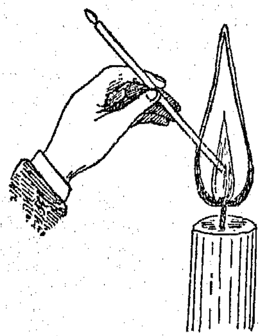


Fig. 43.—Unburnt gases in a candle-flame.

rounding the dark inner zone. In this zone the vapour of the wax is only incompletely burnt, and the solid particles of carbon produced by the decomposition of the wax make it luminous. This zone forms the largest part of the flame.

3. The non-luminous zone (4) near the base where the combustion is complete due to the presence of excess of oxygen of air.

4. There is also a fourth part (3), a faintly luminous zone surrounding the luminous zone (2), but it is hardly visible. This is the hottest part of the flame.

The Bunsen flame (Fig. 42 B) also consists of the same four parts, (3) and (4) being clearly visible.

73. **Ignition temperature.**—Substances take fire at very different temperatures. Some burn at a low temperature, others require a very high temperature. The temperature, to which a substance must be heated before it can burn, is known as its **ignition temperature.** Every combustible substance has its own ignition temperature. Phosphorus, which catches fire in air even at the ordinary temperature, has a low ignition temperature (34°C.). When exposed to air, phosphorus is slowly acted upon by the oxygen of the air and the action is accompanied by the evolution of heat. This heat gradually raises the temperature of phosphorus to its ignition point (34°C.), when it catches fire spontaneously. Carbon bisulphide has a higher ignition temperature than that of phosphorus. It does not burn spontaneously on exposure to air, but is a very inflam-

mable liquid and at once catches fire when a flame is brought into contact with it. Iron, on the other hand, has to be heated to a very high temperature before it can burn in oxygen, and has therefore a high ignition temperature.

The lowest temperature at which the vapour of kerosene oil can take fire when a flame is brought near it, *i.e.*, its ignition temperature, is termed its **flash point.** It varies with the quality of the oil, the cheaper varieties having a low flash point. Kerosene oil having a flash point below a certain temperature (111°F.) is not allowed by Government to be used for burning, as there is a danger of its readily catching fire.

When a substance is burning and its temperature is lowered below its ignition point, it ceases to burn.

Experiment 85.—Place a little methylated spirit in a dish and ignite it. Pour the burning spirit through a wire gauze held on a tripod. As the spirit passes through the gauze, it loses a part of its heat, and the temperature is lowered below its ignition point. It therefore ceases to burn, and may be collected in another dish placed below the gauze.

It is on this principle that **Davy's safety lamp** is constructed (see Heat, page 216).

QUESTIONS.

1. Explain what you understand by the terms: *combustion, combustible,* and *supporter of combustion.* Give two examples of combustion in which oxygen does not take part.

2. What is the luminosity of a flame due to? Give examples to illustrate your answer.

3. Describe the construction of the Bunsen burner.

4. Describe the structure of a candle-flame. How will you show the presence of unburnt gases in the flame ?

5. What is meant by *ignition temperature* and *flash point?* Describe an experiment to show that a combustible substance ceases to burn when the temperature is lowered below its ignition point.

CHAPTER XII.

HYDROCHLORIC ACID AND CHLORINE.

Hydrochloric Acid.

74. **Hydrochloric acid** has long been known in the form of a solution in water, and was named *spirits of salt* or *muriatic acid* on account of its preparation from common salt. In the gaseous form, it is found in small quantities in the gases coming out of the craters of volcanoes.

75. **Preparation.**—(1). Hydrochloric acid gas may be obtained by mixing equal volumes of hydrogen and chlorine in a gas jar (or a bottle) and exposing the mixture to bright sunlight. The two gases combine with an explosive noise and form hydrogen chloride or hydrochloric acid gas.

(2). It is prepared in the laboratory by heating common salt with strong sulphuric acid, when hydrogen chloride is given off and sodium sulphate or Glauber's salt left behind—

Sodium chloride+sulphuric acid
= sodium sulphate+hydrogen chloride.

This method is generally employed in preparing the gas on a commercial scale in Lahore and several other places in the Punjab.

Experiment 86.—Take a flask fitted with a thistle funnel and a delivery tube bent twice at right angles (Fig. 44). Put some common salt in the flask and add sufficient strong sulphuric acid down the funnel to cover the salt. Shake the flask to mix the acid and the salt thoroughly, and warm gently. Hydrochloric acid gas is

evolved and is collected in a jar covered with a perforated cardboard disc by downward displacement of air. Collect a few jars and test the properties of the gas (see below).

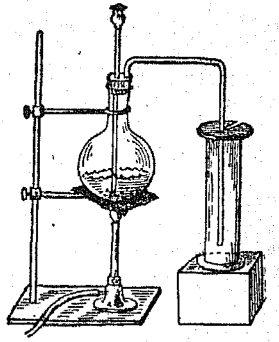


Fig. 44.—Preparation of hydrochloric acid gas.

76. **Properties.** *Physical.* Hydrogen chloride is a colourless gas which fumes strongly in moist air. It has a pungent odour and an acid taste. It is heavier than air and is therefore collected by downward displacement. It is very soluble in water and cannot therefore be collected over it. One volume of water can dissolve about 450 volumes of the gas at the ordinary temperature. The solution obtained by dissolving the gas in water is known as **hydrochloric acid**, a very valuable reagent in the laboratory. The great solubility of the gas in water may be shown by the following experiment:—

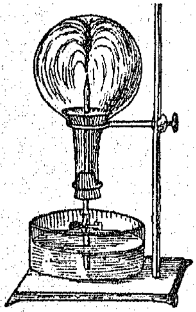


Fig. 45.—Solubility of hydrochloric acid gas in water.

Experiment 87.—*Take a large flask with a tight cork and a glass tube with the* end inside the flask drawn out to a jet. Fit the outer end of the tube with a short piece of india-rubber tubing and close it by means of a pinch-cock. Now remove the cork and fill the flask with the gas. Replace the cork and invert the flask over a trough of water so that the lower end of the tube dips in water (Fig. 45). Open the pinch-cock and

pour a little of ether on the outer surface of the flask to cool it. The gas dissolves in water, and the water rushes up in the flask in the form of a fountain.

Chemical.—The gas is neither combustible nor a supporter of combustion. A lighted candle when introduced in a jar of the gas is extinguished, and the gas does not burn. The gas, when perfectly dry, has no action on litmus, but when dissolved in water, it behaves as a strong acid—*hydrochloric acid.*.

Experiment 88.—Throw a blue litmus paper in a jar of the gas; the paper remains almost unaffected. Moisten the paper and again introduce it in the jar. The paper at once turns red.

When the gas is brought into contact with ammonia, the two gases combine and dense white fumes of ammonium chloride are evolved—

Ammonia+hydrochloric acid=ammonium chloride.

Experiment 89.—Hold a glass rod, with a drop of ammonia solution on it, at the mouth of the gas jar. Dense white fumes are produced. This is used as a test for the gas.

Hydrochloric acid has a sour taste and turns blue litmus red. It readily acts on metals like zinc, magnesium, and iron, forming **chlorides** and giving off hydrogen, *e.g.*—

Zinc+hydrochloric acid
=zinc chloride+hydrogen.

Magnesium+hydrochloric acid
= magnesium chloride+hydrogen.

Experiment 90.—Take some iron filings in a test-tube and add some hydrochloric acid to it. A vigorous action takes place and hydrogen is evolved. Invert another test-tube over the first one and collect the gas. Remove the upper tube after a minute or two and apply a lighted match to it. The gas burns with a slight explosion.

Hydrochloric acid decomposes carbonates, giving carbon dioxide and a chloride (page 84). Sodium carbonate, for example, when treated with hydrochloric acid forms sodium chloride, and carbon dioxide is evolved–

Sodium carbonate+hydrochloric acid
= sodium chloride+water+carbon dioxide.

When treated with caustic soda or potash, it forms sodium or potassium chloride, neutral salt—

Caustic soda+hydrochloric acid
=sodium chloride+water.

The acid, when added to a solution of silver nitrate, gives a white precipitate of silver chloride—

Silver nitrate+hydrochloric acid
=silver chloride+nitric acid.

77. **Tests for hydrochloric acid.**—1. Take a drop of ammonia solution on a glass rod and bring it near the acid. Dense white fumes of ammonium chloride are formed.

2. Add some manganese dioxide to hydrochloric acid in a test-tube and heat gently. A yellowish-green gas, chlorine, is evolved. It has a strong pungent smell and bleaches litmus paper.

3. Dilute the acid with distilled water and add silver nitrate solution. A white precipitate of silver chloride is formed—

Silver nitrate+hydrochloric acid
= silver chloride+nitric acid

Add ammonia to the precipitate; it dissolves. Silver chloride is soluble in ammonia, but insoluble in nitric acid.

78. **Uses of hydrochloric acid.**—Hydrochloric acid is used for making chlorine gas required for the manufacture of chlorides and bleaching powder. It is also used in tanning leather and making dyes.

Mixed with nitric acid, it forms *aqua regia* which is used for dissolving gold and platinum. It is sold in the market generally in stoneware jars.

Chlorine.

79. **History.**—Chlorine was discovered in 1774 by Scheele, who obtained it by heating black oxide of manganese with hydrochloric acid. He believed it to be a compound of hydrochloric acid and oxygen, and named it *oxymuriatic acid*. Some years later, it was proved to be an element and named *chlorine* from its yellowish-green colour.

80. **Occurrence.**—Chlorine is never found free in nature, but in combination with sodium, potassium and magnesium it occurs very abundantly. Its most common compound, sodium chloride or common salt, is found in large quantities in sea-water and salt mines.

81. **Preparation.**—Chlorine is usually prepared in the laboratory by heating manganese dioxide with strong hydrochloric acid, when manganese chloride is formed and chlorine evolved —

Manganese dioxide + hydrochloric acid
= manganese chloride + water + chlorine.

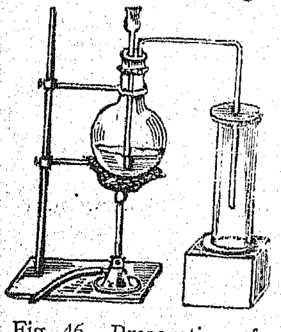


Fig. 46.—Preparation of chlorine.

Experiment 91.—Fit up the apparatus used for the preparation of hydrochloric acid gas (Fig. 46). Place some manganese dioxide in the flask and pour strong hydrochloric acid down the thistle funnel. Heat the mixture and collect the gas in a jar by downward displacement of air. When the jar is full, as seen by the yellowish-green colour of the gas, remove it and quickly

replace it by another. Fill several jars of the gas, and keep each of them covered with a glass disc.

NOTE.—*The gas must be prepared in a fume-cupboard or in open air, as it has a strong pungent odour and is highly poisonous.*

The above method may be modified by using, instead of hydrochloric acid, a mixture of common salt and sulphuric acid, which react together and form hydrochloric acid. This hydrochloric acid further reacts with manganese dioxide and gives chlorine—

Sodium chloride+sulphuric acid
= sodium sulphate+hydrochloric acid.

Manganese dioxide+hydrochloric acid
= manganese chloride+water+chlorine.

This method is sometimes used in preparing the gas on a large scale.

82. **Properties.** *Physical.*—Chlorine is a yellowish-green gas with a strong pungent smell. When inhaled even in a small quantity, it produces cough and cold. It is very poisonous, and if inhaled in large quantities it causes death. It is about $2\frac{1}{2}$ times as heavy as air. It is fairly soluble in water, and cannot therefore be collected over it. The solution of chlorine has the same colour and smell as the gas, and is known as *chlorine water.* The chlorine water is prepared by passing the gas direct from the chlorine apparatus into water.

Chemical.—Chlorine does not burn itself but supports the combustion of several substances. It is a very active gas and combines directly with many elements, such as hydrogen, sodium, phosphorus and antimony, forming **chlorides.**

Experiment 92.—Take a jet of burning hydrogen into a jar of chlorine (Fig. 47). Hydrogen continues to

burn, and white fumes of hydrochloric acid are formed.

Experiment 93.—Place a piece of sodium in a dry deflagrating spoon, and heat the metal till it begins to burn in air. Introduce the burning sodium in a jar of the gas. The metal continues to burn, forming sodium chloride.

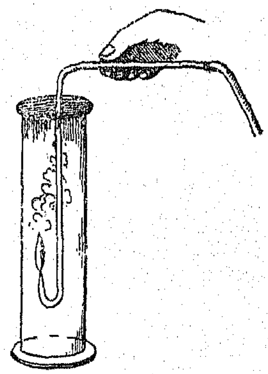


Fig. 47.—Burning of hydrogen in chlorine.

Experiment 94.—Lower a piece of dry phosphorus, held in a deflagrating spoon, in another jar of the gas. The phosphorus takes fire spontaneously, and dense white fumes of chloride of phosphorus are produced.

Experiment 95.—Throw some finely powdered antimony in a jar of the gas. The metal catches fire immediately, and dense white fumes of antimony chloride are formed—

Antimony+chlorine = antimony chloride.

Chlorine has got such a strong chemical attraction for hydrogen that it can combine not only directly with that element, but can also extract it from its compounds, such as water, wax, turpentine oil, etc. If chlorine water be exposed to direct sunlight for a day or two, it loses its colour and smell. The chlorine combines with the hydrogen of water, forming colourless hydrochloric acid and liberating oxygen.

Experiment 96.—Lower a burning candle in a jar of the gas. The candle burns with a dull-red flame, and a good deal of smoke is given off. Add some litmus solution to the jar; it turns red. The wax of the candle is composed of hydrogen and carbon. The chlorine combines with the hydrogen of the candle, forming hydro-

chloric acid and setting free carbon which appears in the form of soot.

Experiment 97.—Soak a piece of dry blotting-paper in hot and pure turpentine and throw it in a jar of chlorine. The turpentine catches fire, and dense clouds of soot are produced. Like candle, turpentine is a compound of carbon and hydrogen. The chlorine combines with hydrogen, and carbon is liberated.

Chlorine is a powerful decolorising and bleaching agent. It can easily bleach vegetable colours, such as litmus, indigo, writing ink and flowers, but it has no action on mineral colours, such as printer's ink. The bleaching action of chlorine is possible only in the presence of moisture. Dry chlorine does not bleach.

Experiment 98.—Throw a few coloured flowers in a jar of chlorine; there is hardly any change in the colour. Now add a little water to the jar and shake; the colour of the flowers disappears.

Repeat the experiment with a piece of printed paper and a paper with a writing in ordinary ink. The writing ink is bleached; while the printer's ink remains unaffected.

The bleaching action of chlorine is due to its strong chemical attraction for hydrogen. Chlorine combines with the hydrogen of water and liberates oxygen—

Chlorine + water

= hydrochloric acid + oxygen

This freshly liberated or **nascent** oxygen, being in a very active state, combines with the colouring matter and forms a colourless compound, and hence the colour is bleached.

83. **Tests for chlorine.**—1. Chlorine gas has a yellowish-green colour and a characteristic pungent smell.

2. It bleaches a moist litmus paper, flower, or a coloured cloth when placed in the gas.

84. **Uses of chlorine.**—Chlorine is frequently employed in bleaching cloth and paper, and in killing disease germs. Chlorine injures silk and straw goods, and cannot therefore be safely used for bleaching these articles. Large quantities of the gas are used in the manufacture of bleaching powder and other disinfectants, and in chlorinating water used for drinking purposes.

85. **Bleaching powder.**—When passed over dry slaked lime, chlorine combines with lime and forms **bleaching powder.** This compound is a yellowish-white powder and readily gives chlorine when treated with a dilute acid, *e.g.*—

Lime chloride (bleaching powder)+hydrochloric acid

= calcium chloride+water+chlorine.

It is for this reason employed as a valuable bleaching agent and disinfectant.

Experiment 99.—Take two beakers, one containing a solution of bleaching powder and the other very dilute sulphuric acid. Dip a piece of coloured cotton cloth first in the bleaching powder solution and then in the acid. Repeat the process till the colour is entirely bleached. Finally wash the cloth with water to remove the excess of the acid.

Bleaching powder is used as a bleaching agent in preference to chlorine, for it is a solid and can be easily transported. It is also used as a disinfectant to destroy germs of various diseases. In cholera days it is dropped in reservoirs for the supply of drinking water in big towns and cities in order to kill the cholera germs.

86. **Chlorides.**—Chlorides are generally prepared—

1. By the direct combination of an element with chlorine. For example, chlorides of hydrogen, phosphorus and antimony are obtained by the direct combination of these elements with chlorine.

2. By the action of hydrochloric acid on metals, their oxides, hydroxides, or carbonates, *e.g.*—

Zinc + hydrochloric acid = zinc chloride + hydrogen.

Zinc oxide + hydrochloric acid = zinc chloride + water.

Zinc hydroxide + hydrochloric acid = zinc chloride + water.

Zinc carbonate + hydrochloric acid = zinc chloride + water + carbon dioxide.

3. Insoluble chlorides, such as silver chloride and lead chloride, are obtained by precipitation—

Silver nitrate+sodium chloride = silver chloride + sodium nitrate.

Lead nitrate + potassium chloride = Lead chloride + potassium nitrate.

Formation of silver chloride as a white precipitate is used as a test for hydrochloric acid or a chloride (page 105).

87. **Tests for chlorides.**—1. Heat a little of the chloride with strong sulphuric acid. White fumes with a strong pungent smell, are given off. Bring a glass rod dipped in ammonia solution in the fumes. Dense white fumes of ammonium chloride are given out.

2. Mix the given chloride with a little of manganese dioxide and heat the mixture with strong sulphuric acid. Chlorine gas is evolved, which is detected by its colour, smell and bleaching action.

3. Shake a little of the chloride with water. If it dissolves, add silver nitrate solution to the solution of the chloride. A white precipitate of silver chloride, soluble in ammonia but insoluble in dilute nitric acid, is obtained.

QUESTIONS.

1. How is hydrochloric acid gas usually prepared and collected? What name is given to the solution of the gas in water?

2. Describe an experiment to show that hydrochloric acid gas is very soluble in water.

3. State the action of dilute hydrochloric acid on (*i*) zinc, (*ii*) washing soda, and (*iii*) silver nitrate.

4. Describe carefully the method of preparation and collection of hydrochloric acid gas. What experiments will you perform to show its properties? Compare it with chlorine. What is its composition? How can it be determined by the synthetic method? [P. U. Matric., 1917.]

5. You are provided with three jars containing oxygen, carbon dioxide and chlorine. What experiments would you perform to distinguish these gases from one another? [P. U. Matric., 1922.]

6. Give the preparation and properties of chlorine gas.

7. How is chlorine obtained from (*i*) common salt, and (*ii*) hydrochloric acid? Enumerate the chief properties of the gas.

8. What elements are present in common salt, bleaching powder, oil of turpentine and candle?

9. Describe what happens, and name the products obtained when a lighted candle is lowered in a jar of (*i*) oxygen, (*ii*) chlorine, and (*iii*) carbon dioxide.

10. What happens when (*i*) a burning candle, (*ii*) a rose flower, (*iii*) powdered antimony, and (*iv*) a filter-paper moistened with hot turpentine are introduced in chlorine? Name the substance formed in each case.

11. Describe experiments to show the bleaching action of chlorine. Why is it necessary to moisten a coloured substance in order that it may be bleached by chlorine?

12. What is bleaching powder? How is it obtained and what are its uses?

CHAPTER XIII.

SULPHUR.

88. **Sulphur** has been known from the earliest times, and was popularly named *brimstone* (fire-stone) on account if its property of burning.

It occurs free (*native*) abundantly in volcanic districts in Sicily, Italy, United States, China, Japan and India. In the Punjab, native sulphur mostly occurs at Jwalamukhi in the Kangra district. In combination with metals, it is found as *sulphides*:—

No.	*Common name.*	*Chemical name.*	*Constituents.*
1.	Iron pyrites.	Iron sulphide.	Iron and sulphur.
2.	Copper pyrites.	Copper sulphide and Iron sulphide.	Copper, ron and sulphur.
3.	Galena.	Lead sulphide.	Lead and su
4.	Zinc blende.	Zinc sulphide.	Zinc and sulphur.

Combined with metals and oxygen, sulphur occurs as *sulphates*:—

No.	*Common name.*	*Chemical name.*	*Constituents*
1.	Gypsum.	Calcium sulphate.	Calcium, sulphur and oxygen.
2.	Heavy spar.	Barium sulphate.	Barium, sulphur and oxygen.
3.	Epsom salts.	Magnesium sulphate.	Magnesium sulphur and oxygen

Sulphur is also present in the white of an egg, and in certain spring waters (page 48).

89. **Extraction.**—Native sulphur, the chief source of sulphur, is always found mixed with earthy impurities, which melt at a much higher temperature than sulphur. To remove these impurities, the native sulphur is piled up in a large kiln with a sloping floor *F* (Fig. 48), and is burnt at

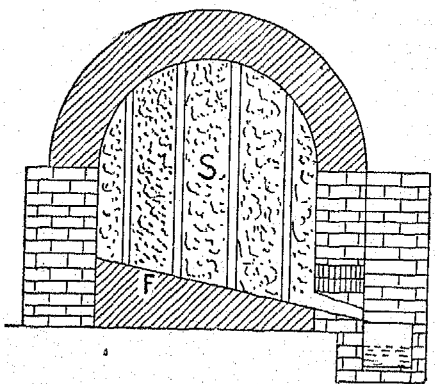


Fig. 48.—Extraction of sulphur.

the bottom. A small quantity burns away, and the heat produced melts the rest of the sulphur which flows down the sloping floor and collects in large troughs, leaving most of the impurities behind.

The crude sulphur thus obtained is purified by distillation. It is placed in an iron pot *P* (Fig. 49), where it is melted by the heat of the furnace *F*. The molten sulphur is then allowed to flow through a pipe into an iron retort *R* placed directly

on the furnace and opening into a large chamber C. When the sulphur in the retort is heated to boiling, the vapour passes into the chamber and condenses on its cold walls in the form of fine crystals, known as **flowers of sulphur**. As distillation proceeds on, the chamber gets heated, and the vapour condenses to the liquid form and collects at the bottom *B*. The liquid is drawn off from an opening *O* into wooden moulds and cast into rods. It is then known as **roll sulphur.**

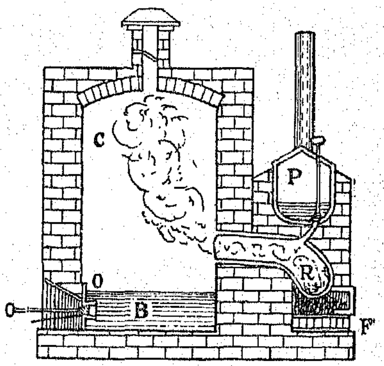


Fig. 49.—Purification of sulphur.

90. **Properties.** *Physical.*—Sulphur is a pale-yellow, brittle solid, which has a faint odour but no taste. It is insoluble in water but soluble in carbon bisulphide. It is about twice as heavy as water. It is a very bad conductor of heat and electricity, and becomes charged with electricity when rubbed with flannel or catskin. It melts at 114°C. to a thin pale-yellow liquid. When heated above its melting-point, it gradually darkens in colour, and becomes thicker and thicker till at a temperature of about 250°C. it is almost black and so thick that the containing vessel may be inverted without the sulphur flowing out. On further heating, the liquid again becomes thin and mobile till at 445°C. it begins to boil, giving off dark-red vapour which may be condensed on a cold surface in the form of a yellow

crystalline powder—*flowers of sulphur*. On cooling, the above changes are reversed.

Experiment 100.—Heat some powdered sulphur in a test-tube very gently, and carefully note the changes in the colour and consistency of the sulphur as the temperature gradually rises. The sulphur at first melts into a thin pale-yellow liquid. Continue heating and see that the liquid becomes thicker and thicker, and finally gets very thick and black in colour. Invert the tube; the sulphur does not flow out of the tube. Heat further and note that the thickness begins to decrease till the liquid boils and gives off dark-red vapour.

91. **Allotropic forms of sulphur.**—Sulphur occurs in three different forms.—

1. Ordinary, rhombic, or octahedral.
2. Monoclinic, prismatic, or needle-shaped.
3. Plastic.

1. **Rhombic sulphur.**—This is the most stable form of sulphur at the ordinary temperature, and other forms, if allowed to stand, gradually pass into this form. Native sulphur and roll sulphur are composed of minute crystals of rhombic sulphur. Large crystals may be easily obtained in the laboratory by the slow evaporation of a solution of sulphur in carbon bisulphide.

Experiment 101.—Take some powdered roll sulphur in a porcelain dish and dissolve it in carbon bisulphide at the ordinary temperature. Filter the solution, if necessary, and allow it to stand in an evaporating dish in the open air. The carbon bisulphide slowly evaporates, and yellow crystals of rhombic sulphur are obtained.

NOTE.—*Carbon bisulphide is a very volatile and inflammable liquid, and no flame should be brought near it.*

2. **Monoclinic or prismatic sulphur.**—This form of sulphur differs from rhombic sulphur in specific

gravity, melting point and the shape of crystals, but resembles it in being soluble in carbon bisulphide. It may be obtained by melting sulphur in a dish and allowing it to cool slowly until a thin crust of solid sulphur appears on the surface. When the crust is broken and the still liquid portion poured off, needle-shaped crystals of monoclinic sulphur are seen in the dish.

Experiment 102.—Take some flowers of sulphur in a dish and heat till the sulphur melts. Allow it to cool until a crust is formed on the surface. Bore a hole in the crust with a round file or a glass rod, and pour out the still liquid sulphur. Now remove the crust; long needle-shaped crystals, which are almost colourless and transparent, will be seen in the dish. Put the dish aside for a day or two; the crystals become opaque and yellow on account of their conversion into the rhombic form. Prismatic sulphur cannot therefore exist unchanged at the ordinary temperature.

3. **Plastic sulphur.**—Unlike the other two forms of sulphur, plastic sulphur is non-crystalline or amorphous, and is insoluble in carbon bisulphide. It is obtained by suddenly cooling sulphur heated nearly to its boiling point.

Experiment 103.—Melt some powdered sulphur in a dish and continue heating till the liquid after becoming thick loses its thickness, and becomes thin and almost black in colour. Pour this dark-coloured liquid into cold water contained in a beaker. A soft, elastic, rubber-like mass—plastic sulphur—is obtained, and can be drawn out into long threads. Allow this plastic mass to stand; it gradually becomes opaque and brittle due to its change into the rhombic form.

92. **Allotropy.**—The above three forms of sulphur, which differ in physical properties, such as colour, density, crystalline form and solubility, are

termed *allotropic modifications* or *allotropes* of sulphur. *Allotropy* means another form, and when the same element exists in two or more forms with different physical properties, the phenomenon is known as **allotropy,** and the different forms are said to be the **allotropic modifications** or **allotropes** of the element. Thus, sulphur exists in three allotropic forms, carbon in three forms—diamond, graphite and amorphous carbon, and phosphorus in two forms—yellow and red. These allotropic modifications are chemically identical and can be converted, one into the other, under proper conditions.

The physical properties of the chief allotropic forms of sulphur may be summarised as below:—

No.	Property.	Rhombic sulphur.	Monoclinic sulphur.	Plastic sulphur.
1.	Physical state.	Rhombic crystals	Needle-shaped crystals.	Amorphous mass.
2.	Specific gravity.	2·04	1·93	1·95
3.	Solubility in carbon di-sulphide.	Soluble	Soluble	Insoluble

93. **Chemical properties of sulphur.**—Sulphur burns in air or oxygen with a bluish flame, forming sulphur dioxide which has a strong suffocating smell like that of a burning match. It combines directly with metals, such as zinc, iron and copper, forming **sulphides**—

Zinc + sulphur = zinc sulphide.
Iron + sulphur = iron sulphide.
Copper + sulphur = copper sulphide.

Experiment 104.—Mix two parts of zinc dust with one part of finely powdered sulphur. Place the mixture on a brick and apply a burning match to it. The mixture

burns with an intense light.

Experiment 105.—Heat some sulphur in a crucible covered with a lid. When the sulphur has melted, remove the lid and add a little of iron filings. A vigorous action takes place accompanied by the production of sparks. Add more iron filings in small quantities at a time. At each addition the mass is seen to glow. Allow the crucible to cool and examine its contents. A greyish-black mass, iron sulphide, is obtained.

Experiment 106.—Take some sulphur in a crucible and heat it to its boiling point. Place a spiral of fine copper wire in the sulphur vapour. The metal becomes red-hot and begins to glow, forming copper sulphide.

94. **Uses of sulphur.**—Sulphur is used in making gunpowder, matches, and in fireworks. It is largely employed in the manufacture of vulcanized rubber, carbon disulphide, sulphur dioxide and sulphuric acid. It is frequently burnt in houses to destroy germs of plague and other infectious diseases.

QUESTIONS.

1. How is pure sulphur obtained on a large scale from native sulphur?
2. Give a short account of the preparation and properties of the different varieties of sulphur.
3. Describe fully the changes that take place when sulphur is heated to its boiling point and then allowed to cool.
4. What is plastic sulphur and how is it obtained? How does it differ from other varieties of sulphur?
5. How would you obtain crystals of sulphur from flowers of sulphur?
6. How would you prove that rhombic, monoclinic and plastic sulphur are chemically identical?

7. What is meant by an *allotropic* variety? Illustrate your answer by referring to two elements which are found in allotropic forms. [P. U. Matric, 1919.]

8. Name some important compounds of sulphur that occur in nature, and give their chemical composition.

CHAPTER XIV.

COMPOUNDS OF SULPHUR.

Sulphur Dioxide.

95. **Sulphur dioxide** is present in the gases coming out from the craters of volcanoes. It is also found in small quantities in the air of large towns, where its presence is due to the combustion of sulphur compounds present in coal.

96. **Preparation.**—(1). Sulphur dioxide is easily prepared by burning sulphur in air or oxygen. This method is largely employed in the preparation of sulphur dioxide required for the manufacture of sulphuric acid.

(2). It is generally prepared in the laboratory by heating copper with strong sulphuric acid. Sulphur dioxide is evolved and copper sulphate left behind—

Copper+sulphuric acid
= copper sulphate+water+sulphur dioxide.

Experiment 107.—Fit up a flask with a thistle-funnel and a delivery-tube (Fig. 50). Put some copper turnings in the flask, and pour strong sulphuric acid down the funnel sufficient to cover the turnings and dip the lower end of the funnel. Heat the mixture and collect the gas evolved by downward displacement of air. Bring a burning match near the mouth of the jar. If it is extinguished, the jar is full of the gas. Remove the jar and quickly replace it by another. Fill a few

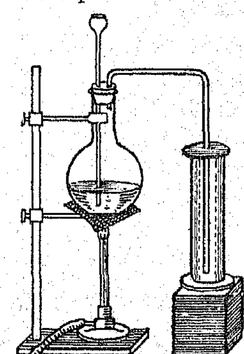


Fig. 50.—Preparation of sulphur dioxide.

jars of the gas and keep each of them covered with a glass disc.

NOTE.—*The gas should be prepared in a fume-cupboard or in open air, as it has a strong suffocating smell.*

97. **Properties.** *Physical.*—Sulphur dioxide is a colourless gas with a strong suffocating smell, characteristic of burning sulphur. It is more than twice as heavy as air. It is very soluble in water, and the solution smells of the gas and has an acid reaction. When this solution is boiled, the whole of the sulphur dioxide is expelled.

Chemical.—Sulphur dioxide neither burns itself nor does it support combustion. If a burning taper be introduced in the gas, the taper is extinguished and the gas does not burn. The aqueous solution of the gas has an acid reaction, because the gas combines with water to form **sulphurous acid**—

Sulphur dioxide+water = sulphurous acid.

When sulphurous acid is treated with an alkali, it forms a **sulphite,** *e.g.* with caustic soda it forms sodium sulphite—

Sulphurous acid+caustic soda
= sodium sulphite+water.

Sulphur dioxide acts as a bleaching agent in the presence of water. The gas takes up oxygen from the water and liberates hydrogen which, being in the nascent state, combines with the colouring matter and reduces it to a colourless compound. The bleaching action of sulphur dioxide is therefore due to reduction, while that of chlorine is due to oxidation (page 109). Moreover, the colour bleached by sulphur dioxide can be regained by exposure to air or by the addition of sulphuric acid, while that bleached by chlorine cannot be restored.

Experiment 108.—Throw a few coloured flowers, moistened with water, in a jar of the gas and shake the jar, or put the flowers in a solution of sulphur dioxide in water; the flowers are bleached. Now add a few drops of sulphuric acid to the jar and shake; the colour is restored.

Sulphur dioxide is a much milder bleaching agent than chlorine and is therefore used in bleaching delicate articles like silk, wool, straw, etc.

When sulphur dioxide is passed through a solution of potassium dichromate, to which some sulphuric acid has been added, the solution turns green. This reaction is used as a test for the gas.

Experiment 109.—Soak a piece of filter-paper in a solution of potassium dichromate acidified with sulphuric acid, and throw the paper in a jar of the gas. The colour changes from yellow to green.

Sulphur dioxide does not combine directly with oxygen, but when a mixture of the two gases is passed over heated *platinised asbestos*, chemical combination takes place and sulphur trioxide is evolved in the form of dense white fumes. Sulphur trioxide readily combines with water, forming sulphuric acid—

Sulphur dioxide+oxygen = sulphur trioxide.

Sulphur trioxide+water = sulphuric acid.

Platinised asbestos is prepared by soaking a piece of asbestos in a solution of platinum chloride and heating strongly. The platinum chloride decomposes and platinum is left as a greyish powder on the asbestos. Platinised asbestos brings about the union of the two gases without itself undergoing any apparent change, and thus it acts as a catalytic agent.

98. Tests for sulphur dioxide.—I. Sulphur

dioxide has a strong suffocating smell like that of a burning match.

2. Introduce a moist litmus paper (or flower) in the gas. The colour is bleached, but it can be restored on dipping the bleached paper (or flower) in dilute sulphuric acid. This distinguishes the gas from chlorine.

3. Dip a filter-paper in a solution of potassium dichromate acidified with a few drops of sulphuric acid. Now place the filter-paper in the gas; it turns green in colour.

99. **Uses of sulphur dioxide.**—Sulphur dioxide is largely used in the manufacture of sulphuric acid and as a bleaching agent for delicate articles like silk, wool, straw and sponge, which would be injured by chlorine. It is a powerful germicide. The disinfecting action of burning sulphur is mainly due to the presence of sulphur dioxide.

Sulphuric Acid.

100. **Sulphuric acid** is one of the most valuable reagents in the laboratory, and is used for many industrial purposes. It is therefore manufactured in very large quantities. On a commercial scale, it is prepared by two methods:—

1. By dissolving sulphur trioxide, obtained by passing a mixture of sulphur dioxide and air over heated platinised asbestos, in water.

2. By passing a mixture of sulphur dioxide, nitric acid, air and steam in large lead chambers, where sulphur dioxide is converted into sulphur trioxide, and the latter dissolves in water to form sulphuric acid.

101. **Properties.** *Physical.*—Sulphuric acid is a thick oily liquid with a specific gravity of 1·85°C. When pure, it is colourless, but the commercial acid has a brown tinge due to the presence of impurities. The pure acid boils at 338°C., but it begins to decompose above 150°C., giving off white fumes of sulphur trioxide. Like other acids, it has a sour taste and turns blue litmus red.

Chemical.—It has a very great affinity for water, and is therefore used as a drying agent for gases on which it has no chemical action. For example, hydrogen, oxygen, air, etc. may be freed from moisture by bubbling them through the acid. When exposed to air, it absorbs moisture and increases both in volume and weight.

Experiment 110.—Weigh out some strong sulphuric acid, say 80 grams, in a beaker, and mark the level of the acid with a gummed paper. Leave the acid exposed to air for a day or two, and note the increase in the volume and the weight of the acid.

When sulphuric acid is mixed with water, a good deal of heat is evolved.

Experiment 111.—Take some water in a beaker and gradually add strong sulphuric acid with constant stirring. The beaker becomes hot. Now place a test-tube containing methylated spirit in the beaker; the spirit begins to boil.

Sulphuric acid must always be diluted by *pouring the acid gradually into water* and constantly stirring the mixture. *Water must never be added to strong sulphuric acid,* because the rapid production of intense heat may cause the liquid to spurt out and produce serious results.

The acid has such a strong affinity for water that it can extract the elements of water from many

organic compounds, such as sugar, wood, paper, etc. These compounds are made up of carbon, hydrogen and oxygen, and are charred when brought into contact with strong sulphuric acid owing to the removal of hydrogen and oxygen and the separation of carbon. This is used as a test for the strong acid.

Experiment 112.—Prepare a thick syrup of sugar in a small beaker. Pour some strong sulphuric acid in the syrup and stir. A brisk action takes place. The mass turns black and swells up to fill the beaker.

Dilute sulphuric acid readily dissolves iron, zinc and magnesium with the formation of **sulphates** and the evolution of hydrogen gas (page 59), but it has no action on copper, lead, or mercury. The concentrated acid, on the other hand, when heated with zinc, copper, or mercury gives sulphur dioxide (expt. 107)—

Copper+sulphuric acid
= copper sulphate+water+sulphur dioxide.

When the acid is heated with chlorides, *e.g.* sodium or potassium chloride, it forms sulphates and gives off hydrochloric acid gas (page 102). It also decomposes carbonates with the evolution of carbon dioxide, *e.g.*—

Sodium carbonate+sulphuric acid
= sodium sulphate+water+carbon dioxide.

Nitrates are also decomposed when heated with the acid, and form sulphates and nitric acid, *e.g.*—

Sodium nitrate+sulphuric acid
= sodium sulphate+nitric acid.

102. **Uses of sulphuric acid.**—Sulphuric acid is the most important of all acids, and is used in the preparation of almost every other acid. It is largely employed in the manufacture of sodium carbonate,

dye-stuffs, and in refining oils. In the laboratory it is frequently used for drying gases. It is also employed in making accumulators used for storing electricity.

103. Sulphates are prepared.—1. By the action of dilute sulphuric acid on metals, their oxides, hydroxides, or carbonates, *e.g.*—

Iron+sulphuric acid
= iron sulphate+hydrogen.

Copper oxide+sulphuric acid
= copper sulphate+water.

Sodium hydroxide+sulphuric acid
= sodium sulphate+water.

Potassium carbonate+sulphuric acid
= potassium sulphate+water+carbon dioxide.

Some metals like copper, lead and silver, are not attacked by the dilute acid, but give the corresponding sulphates when heated with concentrated sulphuric acid, and sulphur dioxide is evolved at the same time, *e.g.*—

Copper+sulphuric acid
= copper sulphate+sulphur dioxide+water.

2. By precipitation.—This method is applied in the case of only those sulphates which are insoluble in water, *e.g.* barium sulphate or lead sulphate. When dilute sulphuric acid or a solution of a sulphate is added to a solution of barium nitrate (or chloride), a white precipitate of barium sulphate is obtained which is insoluble in any acid—

Barium nitrate+sulphuric acid
= barium sulphate+nitric acid.

This is used as a test for sulphuric acid or a sulphate.

The most important sulphates are:—

No.	Common name.	Chemical name.	Constituents.
1.	Glauber's salt. (*khari*).	Sodium sulphate.	Sodium, sulphur and oxygen.
2.	Blue vitriol (*nila tutia*).	Copper sulphate.	Copper, sulphur and oxygen.
3.	Green vitriol (*hira kasis*	Iron sulphate.	Iron, sulphur and oxygen.
4.	White vitriol.	Zinc sulphate.	Zinc, sulphur and oxygen
5.	Gypsum.	Calcium sulphate.	Calcium, sulphur and oxygen.

Calcium sulphate occurs in nature in several forms: (*i*) as **gypsum** in the Madras Presidency, Rajputana and Kalabagh (Punjab); (*ii*) in fine grained condition as **alabaster** in Garhwal (United Provinces), and (*iii*) in transparent crystals as **selenite** in Kathiawar, some parts of the Deccan, and in the marine deposits near Bombay. It is slightly soluble in water and renders the water, in which it is dissolved, permanently hard (page 51). When gypsum is heated to about 130° C., it is converted into a white powder called **plaster of Paris.** This powder when brought into contact with water sets to a hard mass, and is therefore used in making moulds, statues and ornamental vessels.

Experiment 113.—Take a small pill-box, and place a pice or some other coin at its bottom. Prepare a thick paste of plaster of Paris with water. Now fill the box with the paste, and let it stand for some time. When the mass has set, take off the bottom of the box and remove the coin from the hard solid mass. A clear impression of the coin will be seen on the mass.

In the Punjab, plaster of Paris is mostly used in the preparation of black-board chalk. Alabaster

is employed in the manufacture of ornamental cups, saucers and bowls.

104. **Tests for sulphuric acid and sulphates.** Different tests are used for sulphuric acid according as the acid is dilute or strong:—

I. Dilute acid.—

1. In appearance, it is just like water.

2. Take a little of zinc or iron in a test-tube and add some dilute sulphuric acid. Bubbles of hydrogen gas are at once evolved from the surface of the metal.

3. Add a solution of barium nitrate or barium chloride to the acid. A white precipitate of barium sulphate is formed. The precipitate is insoluble in all acids.

II. Concentrated acid.—

1. It is a thick, oily liquid.

2. Dip a match or a piece of wood in the acid. It gets charred and turns black.

3. Heat the acid with copper turnings. A gas, sulphur dioxide, is evolved and is recognised by its smell, resembling that of burning sulphur.

III. Tests for sulphates.—

Most of the sulphates are insoluble in water. No easy test can be applied to insoluble sulphates. Sulphates soluble in water are, however, easily tested by the barium nitrate (or chloride) test.

Experiment 114.—Dissolve a little of sodium sulphate in water in a test-tube. Add barium nitrate (or chloride) solution to the solution of the sulphate. A white precipitate of barium sulphate is formed. Add any acid to this precipitate. It will remain undissolved.

Sodium sulphate + barium nitrate
(soluble) (soluble)
= barium sulphate + sodium nitrate.
(insoluble) (soluble)

QUESTIONS.

1. What happens when copper turnings are heated with concentrated sulphuric acid? How would you proceed to obtain crystals of copper sulphate from the residue left after the reaction?

2. How is sulphur dioxide prepared in the laboratory? How is the gas collected? Why?

3. How would you account for the suffocating smell of a burning match?

4. What would you add to strong sulphuric acid to obtain (*i*) sulphur dioxide, and (*ii*) carbon dioxide?

5. Describe the bleaching action of sulphur dioxide. How does it differ from that of chlorine?

6. Starting with sulphur dioxide, how would you prepare (*i*) sulphurous acid, (*ii*) sulphur trioxide, and (*iii*) sulphuric acid?

7. What is the action of sulphuric acid on (*a*) common salt, (*b*) nitre, (*c*) washing soda, and (*d*) zinc?

8. Describe experiments to show that sulphuric acid has a strong affinity for water.

9. How would you demonstrate the presence of carbon in paper, wood and sugar?

10. What elements are present in galena, gypsum, iron pyrites and blue vitriol?

11. What is plaster of Paris? What is its composition and how is it obtained?

CHAPTER XV.

COMPOUNDS OF NITROGEN.

Ammonia.

105. **History.**—The aqueous solution of ammonia was known to the early chemists, and named *spirits of hartshorn* on account of its preparation by the destructive distillation of hoofs and horns of harts and other animals.

106. **Occurrence.**—Ammonia occurs in small quantities in air and natural waters, where its presence is due to the decay of animal and vegetable matter containing nitrogen. The peculiar pungent smell, often noticed near stables and urinals, is due to the presence of ammonia. In combination, it occurs as ammonium nitrate and ammonium chloride in the soil. A fairly large quantity of ammonium chloride or sal-ammoniac (*noshadar*) is found mixed with earthy impurities near Karnal (Punjab).

107. **Preparation.**—Ammonia is prepared in the laboratory by heating a mixture of ammonium chloride and slaked lime, when ammonia is evolved and calcium chloride remains behind—

Ammonium chloride+slaked lime (calcium hydroxide) = calcium chloride+water+ammonia.

Experiment 115.—Take a mixture of equal quantities of powdered ammonium chloride and slaked lime in a flask fitted with a delivery tube. Gently heat the mixture. Ammonia gas is evolved and collected by upward displacement of air by holding a gas jar in an inverted position over the delivery tube (Fig. 51). Bring a red litmus paper, moistened with water, near the mouth of

the jar. If the paper turns blue, the jar is full of the gas. Fill several jars of the gas and keep them inverted on the table.

2. A large quantity of ammonia is commercially prepared from 'ammoniacal liquors' obtained during the manufacture of coal-gas by the destructive distillation of coal (page 80). The liquor is mixed with slaked lime and heated to give ammonia. The gas is generally passed into hydrochloric or sulphuric acid, by which it is absorbed and forms ammonium chloride or sulphate—

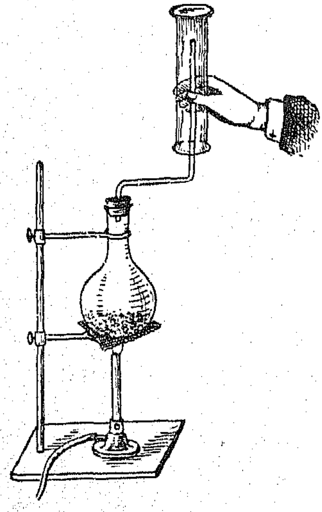


Fig. 51.—Preparation of ammonia.

Ammonia + hydrochloric acid = ammonium chloride.

Ammonia + sulphuric acid=ammonium sulphate.

108. **Properties.** *Physical.* Ammonia is a colourless gas with a strong pungent odour and burning taste. It is lighter than air and extremely soluble in water, and is therefore collected by upward displacement of air. The great solubility of ammonia in water may be shown by the following simple experiment:—

Experiment 116.—Take a jar of the gas and invert it in a trough of water. The water rapidly dissolves the gas and rises in the jar to fill it almost completely.

The solubility of ammonia can also be shown by

the fountain experiment by replacing hydrochloric acid gas by ammonia gas in expt. 87. The solution of ammonia in water is known as *ammonia solution* or simply *ammonia*, and has the smell and properties of the gas. On boiling the solution, the gas is expelled. This gives a convenient method of preparing the gas in the laboratory. When subjected to pressure, ammonia condenses to a colourless liquid even at the ordinary temperature. When liquid ammonia is allowed to evaporate, it absorbs a large amount of heat and produces a considerable cooling. The fact is taken advantage of in ice-machines for making ice.

Chemical.—Ammonia gas does not support combustion, nor does it burn in air. But in oxygen it burns with a yellow flame. It directly combines with acids, forming **ammonium** salts. Thus, when brought into contact with hydrochloric acid, it forms ammonium chloride—

Ammonia+hydrochloric acid
= ammonium chloride.

Experiment 117.—Pour a few drops of strong hydrochloric acid in a jar, and shake the acid so as to moisten the sides of the jar. Invert a jar of ammonia over it, as shown in Fig. 52. Turn the jars upside down so as to mix the ammonia gas with the vapour of hydrochloric acid. Dense white fumes of ammonium chloride are immediately formed.

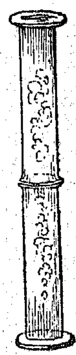


Fig. 52.—Formation of ammonium chloride.

Repeat the experiment with nitric acid; ammonium nitrate is formed.

The formation of ammonium chloride may be

easily shown by holding a glass rod. with a drop of strong hydrochloric acid, over the mouth of a jar of ammonia. Dense white fumes of ammonium chloride are formed. This is used as a test for ammonia or hydrochloric acid gas.

Ammonia solution behaves like an alkali, for it turns red litmus blue and yellow turmeric brown, and combines with acids to form ammonium salts.

109. **Tests for ammonia.**—1. Ammonia has a peculiar pungent smell and excites tears to the eye when inhaled.

2. Hold a moist turmeric paper in the gas. The paper turns brown.

3. Take a drop of hydrochloric acid on a glass rod and hold it in the gas. Dense white fumes of ammonium chloride are evolved.

110. **Uses of ammonia.**—Ammonia is used in ice-making and in the manufacture of washing-soda (sodium carbonate). Ammonium salts are largely employed in medicine. Smelling salts, so often used as a cure for bad cold, contain mostly ammonium carbonate, which readily gives off ammonia on exposure.

Nitric Acid.

111. **Nitric acid** was probably known to the ancient Egyptians, and was termed *aqua fortis* (strong water) on account of its property of dissolving metals. It is one of the common acids and is used in many industries.

112. **Occurrence.**—Nitric acid occurs free only in minute traces in the atmosphere, but in the form of nitrates it is widely distributed in nature. Sodium nitrate or *Chili saltpetre* (containing sodium, nitro-

gen and oxygen) occurs in large deposits in Chili and Peru. Potassium nitrate (containing potassium, nitrogen and oxygen), commonly known as *salt-petre* or *nitre* (*shora*), is found in several parts of India. In the Punjab it is chiefly found in Shahpur, Mianwali and Rohtak districts.

113. **Preparation.**—Nitric acid is prepared by heating nitrate of sodium or potassium with strong sulphuric acid, when nitric acid is evolved and sodium or potassium sulphate left behind—

Sodium nitrate+sulphuric acid
= sodium sulphate+nitric acid.

Experiment 118.—Put some nitre in a glass retort and place the retort on a sand-bath (an iron tray containing sand) or on a wire-gauze supported on a tripod stand. Fix the retort in a clamp, slip a flask over its neck, and support the flask in a trough of water (Fig. 53). Pour strong sulphuric acid in the retort by means of a funnel till the nitre is covered. Stopper the retort and heat the mixture gently. Reddish-brown fumes of nitric acid are evolved and condense to a pale-yellow liquid in the flask which is cooled by a regular stream of water. When sufficient nitric acid has been collected in the receiver, disconnect the apparatus and test the properties of the liquid (see below).

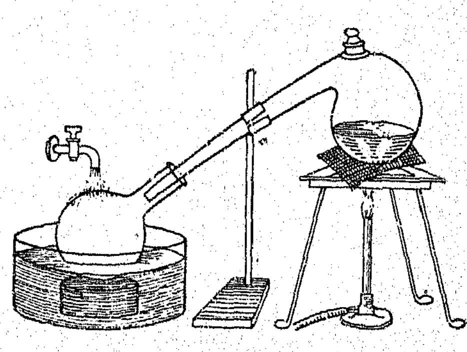


Fig. 53.—Preparation of nitric acid.

This method is employed in the preparation of nitric acid on a manufacturing scale. Nitre or Chili saltpetre and sulphuric aiid are heated in cast iron retorts and the acid vapour condensed in earthenware pipes cooled by water, and collected in stoneware jars.

114. **Properties.** *Physical.*—Pure nitric acid is a colourless liquid with a peculiar pungent odour. The commercial acid is generally pale-yellow in colour on account of the presence of oxides of nitrogen produced by its slight decomposition during distillation. It fumes strongly in moist air. It is highly corrosive and causes painful wounds when brought into contact with the skin. It produces yellow stains on organic substances, *e.g.* skin, wood, cloth, paper, etc. The pure acid has a specific gravity of about 1·5 and boils at 86°C.

Chemical.—Nitric acid readily decomposes on heating, giving off oxygen and oxides of nitrogen—

Nitric acid=water+nitrogen peroxide+oxygen.

Experiment 119.—Take a clay-pipe with a long stem and fix it in a clamp in a slanting position (Fig. 54). Let the end of the stem dip in water in a pneumatic trough. Invert a test-tube full of water over the end, and heat the middle of the stem. When the stem is red-hot, pour some concentrated nitric acid, drop by drop, into the bowl of the pipe. As the acid passes down the heated part of the stem, it decomposes to give off oxygen gas which is

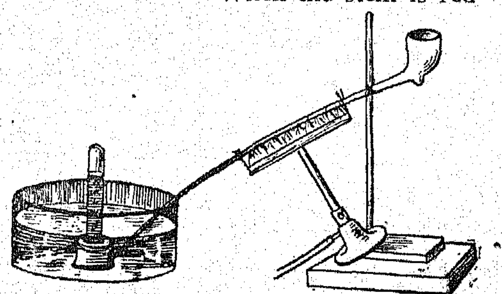


Fig. 54.—Decomposition of nitric acid.

collected in the test-tube. When sufficient gas has been collected in the tube, stop heating and test the gas for oxygen with a glowing splinter of wood.

On account of the ease with which it gives up oxygen, nitric acid acts as a powerful oxidising agent. Carbon and sulphur are oxidised to carbon dioxide and sulphuric acid respectively when heated with the acid.

Experiment 120.—Place some glowing pieces of charcoal on a stone slab, and carefully pour a little of strong nitric acid on them. A vigorous action takes place due to the rapid oxidation of the charcoal by the acid. The charcoal bursts into flame, and reddish-brown fumes of nitrogen peroxide are evolved.

When the acid is dropped on heated sawdust, the oxidation is so rapid that the mass catches fire.

Experiment 121.—Heat some fine sawdust in an iron tray till it just begins to char. Pour a few drops of strong nitric acid on the heated mass. The sawdust at once takes fire and reddish-brown fumes of nitrogen peroxide are given off.

Most metals are attacked by nitric acid, forming **nitrates.** Unlike other acids, hydrogen is very seldom evolved when nitric acid acts on metals. Copper, which is unaffected by hydrochloric or cold sulphuric acid, is readily acted upon by nitric acid, forming copper nitrate and giving off reddish-brown fumes of nitrogen peroxide—

Copper+nitric acid

= copper nitrate+nitrogen peroxide+water.

Experiment 122.—Pour a little nitric acid on some copper turnings contained in a porcelain dish. The liquid turns bluish-green due to the formation of copper nitrate, and reddish-brown fumes of nitrogen peroxide are evolved.

Gold and platinum do not dissolve in this acid or in any other single acid, but they are soluble in a

mixture of nitric and hydrochloric acids. This mixture is known as *aqua regia* (royal water), for it can dissolve gold, the king of metals.

115. **Uses of nitric acid.**—Nitric acid is used in extracting gold and silver from laces, in etching designs on copper vessels, and in the preparation of nitrates. It is largely employed in the manufacture of sulphuric acid, dye-stuffs and explosives.

116. **Nitrates.**—The presence of nitrates in nature has already been mentioned. They are formed in nature by the decomposition of nitrogenous organic matter. Ammonia and its compounds, formed during the decomposition, are gradually oxidised to the nitrates through the agency of *nitrifying bacteria* present in the soil.

Nitrates are generally prepared by the action of nitric acid on metals, and are soluble in water. Thus, copper and silver nitrates are prepared by dissolving the metals in the acid. The nitrates may also be prepared by the neutralization of nitric acid with bases. For instance, sodium (or potassium) nitrate may be obtained by adding a sufficient quantity of caustic soda (or caustic potash) solution to neutralize the acid—

Caustic soda+nitric acid
= sodium nitrate+water.

All nitrates decompose more or less readily on heating, giving off oxygen, and in many cases oxides of nitrogen at the same time.

Experiment 123.—Heat strongly some nitre in a test-tube; a crackling noise is produced. Introduce a glowing splinter in the gas evolved. The splinter bursts into flame, showing the presence of oxygen.

Repeat the experiment with lead nitrate and see that it decomposes to give off oxygen and reddish-brown fumes

of nitrogen peroxide—

Lead nitrate = lead oxide + nitrogen peroxide + oxygen.

On heating a nitrate with strong sulphuric acid, nitric acid is evolved and can be distinguished by its peculiar pungent odour. If some copper turnings be added to the mixture, reddish-brown fumes of nitrogen peroxide are given off, and the liquid turns bluish-green due to the formation of copper nitrate. This reaction can be used as a test for nitrates and nitric acid.

Experiment 124.—Add a few copper turnings to nitric acid or to a mixture of a nitrate and strong sulphuric acid contained in a test-tube. Heat the tube gently. Reddish-brown fumes of nitrogen peroxide are evolved and the liquid turns bluish-green in colour.

Nitrates can also be identified by the 'ring test.'

Experiment 125.—Take a solution of nitre in water in a test-tube, pour strong sulphuric acid into the solution and cool the tube. Now add a little ferrous sulphate solution along the sides of the tube. A dark-brown ring is formed at the place where the two liquids meet.

Sodium and potassium nitrates are used in the manufacture of nitric acid, in fire-works, and as fertilizers. Silver nitrate is a valuable reagent in the laboratory, and is also used in photography.

117. **Distinguishing tests of the three acids.**—The distinguishing tests of hydrochloric, sulphuric and nitric acids are given in the table on the next two pages.

DISTINGUISHING TESTS OF HYDROCHLORIC, SULPHURIC AND NITRIC ACIDS.

Concentrated acids:—

No.	Experiment.	Hydrochloric acid.	Sulphuric acid.	Nitric acid.
1.	Examine the colour.	When pure, colourless. The commerical acid is generally bright yellow due to the presence of chlorides of iron.	When pure, colourless oily liquid. The commercial acid has a brownish tinge due to the presence of oxides of nitrogen.	When pure, colourless. The commercial acid is straw-yellow due to the presence of nitrogen peroxide.
2.	Notice the nature of fumes when exposed to air.	Fumes evolved in moist air have a characteristic irritating odour.	No fumes.	Fumes evolved have a pungent odour.
3.	Hold a drop of ammonia solution on a glass rod over the liquid.	Dense white fumes of ammonium chloride are evolved.	No fumes.	White fumes of ammonium nitrate are formed.
	Drop a little of the liquid on a piece of paper or wool, and notice the colour.	Yellow tinge. The stains on a black cloth are at first red but then change to reddish-brown.	Charred. When dropped on a black cloth, a brown stain is formed and the fibre is destroyed.	Reddish-brown. The colour is darkened by the action of caustic alkalies. Coloured clothes are corroded and stained yellow by the acid.
5.	Add a little of copper to the acid in a test-tube.	No reaction even on boiling.	Reacts only on heating, giving off suffocating odour of sulphur dioxide like that of a burning match.	Violent action with the evolution of reddish-brown fumes of nitrogen peroxide.

Concentrated acids (continued).—

Confirmatory tests.	When heated with manganese dioxide, chlorine gas is evolved. The gas is recognised by the yellowish-green colour, irritating smell and bleaching action on moist litmus paper.	A filter-paper, soaked in a solution of potassium dichromate and dilute hydrochloric acid, held in the gas evolved by heating the said with copper, turns green.	1. A dilute solution of indigo is bleached by the acid. 2. When poured on a red-hot charcoal, a vigorous reaction takes place due to the rapid oxidation of the charcoal by the acid. 3. Gives the 'Ring test' (see below).

Dilute acids.

No.	Experiment.	Hydrochloric acid.	Sulphuric acid.	Nitric acid.
1.	Add a little of silver nitrate solution to the liquid in a test-tube.	A white curdy precipitate, insoluble in nitric acid but soluble in ammonia solution, is obtained.	No precipitate, but only a milkiness.	No precipitate.
2.	Add a little barium chloride solution to the liquid in a test-tube.	No precipitate.	White precipitate of barium sulphate, insoluble in any acid, is formed.	No precipitate. Confirmatory test.— When thoroughly mixed with strong sulphuric acid and then a little ferrous sulphate solution is added gently along the sides of a cooled test-tube, a dark-brown ring is formed where two liquids meet. The ring disappears on shaking or warming the tube. This is known as the '**ring test**.'

QUESTIONS.

1. How would you prepare and collect ammonia gas? Make a drawing of the apparatus used.

2. Describe an experiment to show that ammonia is very soluble in water.

3. Give the properties and uses of ammonia gas.

4. What are the chief sources of ammonia? How are ammonium compounds prepared from the gas?

5. How is ammonia solution prepared and how is the ammonia gas obtained from it?

6. Compare the properties of a solution of ammonia with those of caustic soda and slaked lime.

7. Describe the preparation of nitric acid, and explain its action on metals.

8. What is meant by an oxidising agent? Describe experiments to show the oxidising action of nitric acid.

9. Give the characteristic tests of nitric acid and nitrates.

10. Describe the preparation of nitric acid, and give experiments to show the oxidising properties of this acid and its salts. [P. U. Matric., 1918.]

11. You are provided with four bottles containing water and solutions of sulphuric acid, hydrochloric acid and ammonia. What tests would you try to distinguish them from one another? [P. U. Matric., 1923.]

CHAPTER XVI.

PHOSPHORUS AND ITS COMPOUNDS.

Phosphorus.

118. **History.**—Brand, a German alchemist, was in search of the *philosopher's stone*, a substance which would convert base metals into gold. He happened to distil concentrated urine with sand and accidentally discovered phosphorus. The word 'phosphorus' means light-bearer, and the element was so named on account of its peculiar property of glowing in the dark.

119. **Occurrence.**—Phosphorus does not occur free in nature, as it readily combines with the oxygen of the air. In the combined state it is present in urine, blood, brain, bones and other parts of the body. It also exists in plants, seeds and fruits, from which it passes to the animals that feed on them. In fact, phosphorus is essential for the growth and development of both animals and plants. In combination with metals and oxygen, it occurs as **phosphates** which are very widely distributed in nature. Calcium phosphate, a compound of calcium, phosphorus and oxygen, is present in all fertile soils, and is the chief constituent of bones.

120. **Preparation.**—Phosphorus is mostly obtained from *bone-ash,* a residue left after burning bones, which contains nearly 90 per cent. of calcium phosphate by two different methods: (*i*) retort process, (*ii*) electrical process.

1. *Retort process.*—Bone-ash is first treated with sulphuric acid in large wooden vessels, when

calcium sulphate and phosphoric acid are produced—

Calcium phosphate+sulphuric acid
=calcium sulphate+phosphoric acid.

The phosphoric acid is then separated from the insoluble calcium sulphate by filtration and concentrated to a thick syrup. The syrup is next mixed with powdered charcoal and heated in an iron vessel to drive out water from the phosphoric acid. The dried mass is then heated to a white-heat in clay retorts, when phosphorus distils over and is condensed under water. The phosphorus thus obtained is melted under water, purified, and cast into sticks.

This method has now been replaced by the electrical process.

2. *Electrical process.*—Phosphorus is almost entirely prepared by the electrical process. Here bone-ash is mixed with sand (silica) and coke, and the mixture is strongly heated in an electric furnace with carbon electrodes (Fig. 55). Phosphorus vapour passes out through an opening O and is condensed under water. A fresh charge of the mixture is added from H, and the process is thus made continuous.

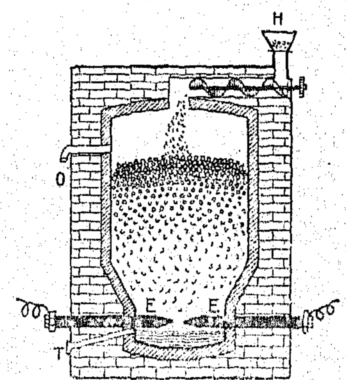


Fig. 55.—Manufacture of phosphorus by electric furnace.

121. **Properties.**—Phosphorus exists in two

allotropic forms: (*i*) yellow or ordinary phosphorus, and (*ii*) red phosphorus.

1. **Yellow phosphorus.**—This form of phosphorus, which is obtained by the method above described, is a pale-yellow, translucent solid. It is soft like wax at the ordinary temperature and can be cut with a knife but on exposure to light it gradually becomes brittle. It fumes in moist air and catches fire on exposure, and is therefore kept under water. It has a specific gravity of 1·8, and melts at 44°C. when heated under water. It is insoluble in water, but readily soluble in carbon bisulphide. It is very poisonous, and workers in match factories, where yellow phosphorus is used, generally suffer from diseases of the jaw and nose owing to their frequent exposure to the fumes of phosphorus.

Yellow phosphorus is a highly inflammable substance. It catches fire at 34°C. and gives dense white fumes with a garlic (*lasan* or *thom*) smell. When exposed to air, phosphorus combines slowly with oxygen, giving off white fumes of phosphorus pentoxide. If the room, in which phosphorus is exposed, be sufficiently warm, or if the exposure be sufficiently long, the element may even catch fire.

In a finely divided state phosphorus is oxidised so readily that it ignites spontaneously.

Experiment 126.—Cut two or three pieces of yellow phosphorus under water and dry them carefully between the folds of a filter-paper. Dissolve the pieces in carbon disulphide and pour a little of the solution on a filter-paper placed on an iron tray. The carbon disulphide evaporates rapidly and leaves the phosphorus behind in such a fine state that it catches fire spontaneously due to rapid oxidation by air.

Experiment 127.—Cut one or two pieces of phosphorus under water and carefully dry them between the folds of a filter-paper. Dissolve the pieces in carbon disulphide, and pour a little of the solution over a filter-paper held by a pair of tongs. The carbon disulphide rapidly evaporates and leaves the phosphorus in such a finely divided state that it takes fire spontaneously due to rapid oxidation by air.

NOTE.—*Phosphorus is highly inflammable, and must not be touched with the fingers. It must always be cut under water and handled with a pair of forceps.*

The oxidation of phosphorus is accompanied by the emission of a faint yellow light which is distinctly visible in the dark. The phenomenon is known as *phosphorescence.*

Experiment 128.—Place a few pieces of yellow phosphorus in a little water contained in a flask. Fit the flask with a wide glass tube, about a foot long. Boil the water in the flask. The steam evolved carries with it some phosphorus vapour which produces a feebly luminous ring in the tube at the point where the steam condenses. The ring is best seen in a dark room.

2. **Red or amorphous phosphorus**—This is a dark-red, brittle powder having a specific gravity of 2·1. It is obtained by heating yellow phosphorus to about 250°C. in closed vessels out of contact with air. While yellow phosphorus, when exposed to air or light, gradually passes into the red variety, the red phosphorus remains unchanged for any length of time, and is therefore the stable form of phosphorus at the ordinary temperature. When heated, it directly passes into vapour without melting. It is non-poisonous, and is insoluble in water and carbon bisulphide. It does not fume in air, nor does it emit any light in the dark. It catches fire only when heated to about 260°C. in air, and is therefore

much less reactive than the yellow phosphorus.

Experiment 129.—Place a little of yellow and red phosphorus side by side on an iron tray placed on a tripod and gently heat over a small flame. The yellow phosphorus soon catches fire and burns brilliantly, whereas the red phosphorus remains unaffected.

Both the forms of phosphorus yield the same oxide (phosphorus pentoxide) on burning in air, and are therefore chemically identical.

122. **Yellow and red phosphorus.**—The principal differences in the properties of the two varities of phosphorus are summarised in the following table:—

	Yellow phosphorus.	Red phosphorus.
1.	A pale-yellow, semi-transparent solid, gradually turning brown, and finally almost black on exposure to light.	A dark-red opaque solid, unchanged by exposure to light, and hence regarded as the stable form of phosphorus at the ordinary temperature.
2.	Soft like wax.	Brittle powder.
3.	Very poisonous	Non-poisonous.
4.	Density 1.	Density 2·1.
5.	Melts at 44°C. when heated under water.	Does not melt, but directly changes to vapour when heated to 250°C.
6.	Emits a faint yellow light in the dark—*Phosphoresces.*	Emits no light.
7.	Fumes in air and takes fire when exposed to air or heated to 34°C. in air, hence kept under water.	May be kept unchanged in air for any length of time at the ordinary temperature, and takes fire only when heated in air to above 260°C.
.	Readily soluble in carbon bisulphide.	Insoluble in carbon bisulphide.

123. **Uses of phosphorus.**—Yellow phosphorus is largely employed in the manufacture of ordinary or lucifer matches. The tip of an ordinary match is coated with a mixture of yellow phosphorus, antimony sulphide, potassium chlorate and glue made into a thick paste with water. When this tip is rubbed against a rough surface, the heat caused by friction ignites the phosphorus, and the match burns.

On account of the strongly poisonous nature of yellow phosphorus, the use of ordinary matches has been restricted in some countries by law, and 'safety' matches are now commonly used. The head of a safety match contains no phosphorus, but is composed of antimony sulphide, potassium chlorate and glue, while the surface on which the match is rubbed for ignition is composed of red phosphorus, antimony sulphide and powdered glass.

Yellow phosphorus is also used in small quantities in medicine and in making rat-paste which is composed of yellow phosphorus, lard and flour made into a paste.

Compounds of Phosphorus.

124. **Oxides.**—Phosphorus forms two important oxides: phosphorus trioxide and phosphorus pentoxide.

1. **Phosphorus trioxide.**—This oxide is formed mixed with the pentoxide when phosphorus is exposed to air or burnt in a limited supply of air. It is a white crystalline solid having an unpleasant smell. When heated in air, it takes fire and forms phosphorus pentoxide—

Phosphorus trioxide+oxygen
= phosphorus pentoxide.

Phosphorus trioxide slowly combines with cold water to form **phosphorous acid.** The phosphorous acid, when treated with alkalies, gives salts known as **phosphites.** Thus, with caustic soda it forms sodium phosphite—

Phosphorous acid+caustic soda
= sodium phosphite+water.

Both phosphorus trioxide and phosphorous acid are poisonous.

2. **Phosphorus pentoxide.**—This oxide is obtained by burning phosphorus in an excess of air or oxygen (see expt. 38). It is a colourless, amorphous powder which, when pure, has got no smell. It fumes strongly in moist air, and combines with water with a hissing noise to form **phosphoric** acid. On account of its strong affinity for water, it acts as a powerful *dehydrating agent*—a substance that can remove the elements of water from other substances. For instance, when brought into contact with sulphuric acid, it takes up water from the acid and forms phosphoric acid, while sulphuric acid is converted into sulphur trioxide—

Phosphorus pentoxide+sulphuric acid
= phosphoric acid+sulphur trioxide.

For the same reason it chars paper, wood and several other organic substances. It is frequently used as a strong drying agent in the laboratory.

125. **Phosphoric acid,** as stated above, is formed when phosphorus pentoxide dissolves in water. It may also be obtained by heating phosphorus with concentrated nitric acid. Nitric acid is a strong oxidising agent (page 136) and oxidises phospho-

rus to phosphoric acid. Phosphoric acid is prepared on a large scale by treating bone-ash with sulphuric acid (page 143). It is a colourless crystalline solid which combines with alkalies to form phosphates, *e.g.*—

Caustic soda+phosphoric acid
= sodium phosphate+water.

QUESTIONS.

1. How is phosphorus prepared? Describe its properties and uses.

2. In what respects does yellow phosphorus differ from red phosphorus? How would you show that they are allotropic forms?

3. Name the products obtained when phosphorus burns in a limited supply of (*i*) air, (*ii*) oxygen, and (*iii*) chlorine.

4. How do the ordinary matches differ from the safety matches? Why does an ordinary match ignite when rubbed on a rough surface?

5. Distinguish between the properties of red and yellow phosphorus, and mention what you know about phosphorus pentoxide. [P. U. Matric., 1915.]

6. How is phosphoric acid obtained from (*i*) phosphorus, and (*ii*) bone-ash?

7. Give as many examples as you can of simple substances which differ in appearance and physical properties, but are identical in chemical composition. State briefly their principal characteristics. [P. U. Matric., 1924.]

PANJAB UNIVERSITY QUESTION PAPERS.

Chemistry—Non-Metals.

1927

1. What do you understand by a physical change and a chemical change? State with reasons whether the changes taking place during the following processes are chemical or merely physical :—
   (*a*) crystallisation, (*b*) sublimation, (*c*) respiration, (*d*) combustion, (*e*) electrolysis, and (*f*) rusting.

2. Describe experiments to show that oxygen, nitrogen, carbon dioxide and water vapour are the principal constituents of air. In what proportions are these gases present in the atmosphere?

3. Describe experiments to show the action of—
   (*a*) hydrochloric acid on copper,
   (*b*) sulphuric acid on lead,
   (*c*) nitric acid on mercury, and
   (*d*) plants on carbon dioxide.

   State clearly the conditions under which the above reactions would take place and the products of reaction formed in every case.

4. In what respects do spring-water, mineral water, sea-water and rain-water differ from one another? How would you prepare fresh water from a dirty sample of sea-water?

5. Describe the preparation, properties and uses of oxygen.

6. How is ammonia prepared? Give a short account of its chief properties and uses.

7. Describe what you know about the different varieties of sulphur. In what form does sulphur occur in nature? Describe carefully the action of heat on sulphur.

1928

1. Describe the method you would employ to prepare a pure specimen of the following :—
   (*a*) copper sulphate from commercial blue vitriol,
   (*b*) ammonium chloride from a mixture of ammonium chloride and sodium chloride,
   (*c*) oxygen from a mixture of oxygen and carbon dioxide,
   (*d*) sand from a mixture of sand and zinc dust.

2. Describe very carefully the changes (both physical and chemical) that take place when excess of (*a*) sulphur, and (*b*) magnesium are heated in a small jar of air. Prove your results by means of simple experiments.

3. 'The air is not a chemical compound but a mixture of gases.' What experimental proofs have you in support of it?

4. Write an account of the preparation and properties of hydrochloric acid gas.

Or,

How would you prepare the following?
(*a*) nitrates, (*b*) chlorates, and (*c*) sulphates.
What are their characteristic uses? Give examples.

5. Write short notes on the following:—
(*a*) allotropy, (*b*) synthesis, (*c*) water of crystallisation, (*d*) electrolysis, (*e*) precipitation, and (*f*) amalgams.

1929

1. Describe the methods you would employ to prepare a pure specimen of the following:—
(*a*) drinking water from sea-water,
(*b*) mercury chloride from potassium chloride,
(*c*) sulphur from iron pyrites,
(*d*) sulphur from a mixture of sulphur and sand.

2. Describe experiments to show the exact quantities of solids and gases that are present in sea-water.

Or,

What will happen if you heat the following substances in a closed jar containing nitrogen obtained from air:—
(*a*) red oxide of mercury, (*b*) potassium chlorate, (*c*) magnesium, and (*d*) phosphorus?

3. What tests would you employ to distinguish the dilute solutions of sulphuric acid, nitric acid, and hydrochloric acid from one another?

Or,

How would you prepare the following:—
(*a*) nitrates, (*b*) chlorates, and (*c*) sulphates?
What are their distinguishing tests?

4. A mixture of hydrogen and hydrochloric acid gas is given to you. How would you separate the gases, and what characteristic tests would you perform to distinguish one from the other?

1930

1. Explain clearly, with the help of experiments, what you understand by—

(*a*) chemical combination and decomposition, and (*b*) oxidation and reduction.

2. What do you understand by a solution? How would you prepare a saturated solution at a given temperature? What is generally the effect of cooling such a solution? In what important respects does a solution differ from an ordinary mixture?

3. Write what you know about the composition of air. Describe experiments to show the changes brought about in air by—

(*a*) the burning of a candle, and (*b*) the breathing of animals.

4. Describe experiments to show what happens when—

(*a*) powdered coal is heated strongly in the absence of air,
(*b*) phosphorus burns in a free supply of air, and
(*c*) sulphur is heated in air and in a closed vessel.

Or,

Write what you know about the action of—

(*a*) sulphuric acid on common salt,
(*b*) hydrochloric acid on marble,
(*c*) caustic soda on aluminium,
(*d*) heat on a mixture of ammonium chloride and slaked lime.

What tests would you try to identify the gas evolved in each case?

5. Write a short account of the different forms of the element, carbon. By what experiments would you prove that they are really different forms of the same element?

1931

1. What do you understand by the hardness of water? What are the different kinds of hardness and their causes? What happens when hard water is used for washing and boiling purposes? Is there any objection to the use of hard water for drinking purposes?

2. Describe experiments to show—

(*a*) the oxidising action of nitric acid,
(*b*) the great solubility of hydrochloric acid gas,
(*c*) the bleaching action of sulphur dioxide,
(*d*) the great affinity of sulphuric acid for water.

Or,

Describe, in detail, the method that you would employ to separate the constituents of—

(*a*) finely powdered chalk and water,

(*b*) ammonium chloride and sand,
(*c*) sulphur and powdered charcoal,
(*d*) nitre and common salt.

3. Describe, with full practical details, the most convenient method of preparing oxygen in the laboratory. What experiments would you try to demonstrate that oxygen is a great supporter of combustion? Write what you know about the important methods used for obtaining the gas on a large scale.

4. Write what you know about the allotropic varieties of phosphorus and the characteristic differences between them. How would you show their identity?

What is the principal source of phosphorus and what are its applications?

1932

1. Describe an experiment by which you can prove the indestructibility of matter.

2. What is the difference between a chemical compound and a mechanical mixture? How would you distinguish one from the other?

3. How would you establish the composition of water?

Or,

How would you prepare ammonia in the laboratory? What experiments would you perform to illustrate its properties?

4. Write short notes on :—
(*a*) sublimation, (*b*) amalgams, and (*c*) allotropy.

Or,

How will you find whether a bottle contains sulphuric, nitric, or hydrochloric acid?

5. In how many different forms can sulphur exist? In what ways do the different forms differ from each other?

What are the principal uses of sulphur? How would you purify an impure specimen of sulphur?

1933

1. Describe the preparation of ammonia. Describe three experiments to illustrate its properties.

Or,

How is hydrogen obtained in the laboratory? Write what you know about its important properties and uses.

2. Describe what happens when (*a*) powdered antimony, (*b*) moist red flower, and (*c*) a filter-paper soaked in turpentine are introduced into a jar of chlorine. How is chlorine prepared?

Or,

How is bleaching powder obtained? What are its uses?

3. Describe what happens when concentrated sulphuric acid acts on (*a*) common salt, (*b*) nitre, (*c*) sugar, (*d*) lead (*e*) copper, and (*f*) phosphorus pentoxide.

4. What experiments will you try to find out whether a sample of water is hard or soft? Explain what are the different kinds of hardness.

5. In what form does phosphorus occur in nature, and how is it extracted?

Or,

How would you prove that yellow and red phosphorus are chemically identical?

1934

1. Describe the preparation of oxygen gas. Describe three experiments to illustrate its properties.

Or,

What are the products formed when (*a*) heated iron is introduced into a jar of oxygen, (*b*) a jar of hydrogen is mixed with a jar of oxygen, (*c*) phosphorus and sulphur are burnt in oxygen? How would you test for the products formed in each case?

2. Describe the preparation, properties and uses of nitrogen.

Or,

What happens when (*a*) silver, (*b*) lead, (*c*) copper, and (*d*) ammonia are acted upon by dilute nitric acid?

3. How would you prepare a sample of pure water from the following :—

(*a*) dilute solution of ammonia,
(*b*) dilute solution of nitric acid,
(*c*) sea-water.

How would you test for the purity of water in each case?

4. What is allotropy? In what forms does sulphur occur?

Or,

From a mixture of nitrogen and hydrogen, how would you prepare pure nitrogen?

1935

1. Describe the composition of air. How would you demonstrate the presence of any two constituents of air?

Or,

What do you understand by the terms: *element* and *compound*?

Classify the following substances into elements and compounds :—

(*i*) lime, (*ii*) iron filings, (*iii*) sulphur, (*iv*) diamond, and (*v*) common salt.

2. You are provided with three jars containing ammonia, hydrogen chloride and carbon dioxide. How would you prove which is which?

Or,

How is ammonia prepared in the laboratory? Describe three experiments to illustrate its properties.

3. What experiments would you try to find out whether a sample of water is hard or soft? What are the different kinds of hardness, and what are the causes of hardness?

Or,

What do you understand by *oxidation* and *reduction*. Give examples of oxidising and reducing agents.

4. What are the different forms of carbon? How would you prove that they are really the different forms of the same element?

Or,

Describe the preparation and properties of sulphur dioxide.

1936

1. What is meant by a physical change and a chemical change? How would you classify the following, giving your reasons in each case :—

(*a*) Glowing of an electric bulb.
(*b*) Burning of a candle.
(*c*) Rusting of iron.
(*d*) Melting of ice to water.

Or,

What do you understand by the terms: *solution, sublimation* and *distillation*? Give experiments to illustrate the uses of these operations in chemistry.

2. Describe fully how oxygen is prepared in the laboratory and on a large scale. Give *three* experiments to illustrate its properties.

Or,

Describe th- preparation, properties and uses of hydrogen.

3. How would you determine the composition of water?

Or,

Describe fully what happens when—

(*a*) Copper sulphate crystals are heated in a test-tube.
(*b*) Aluminium powder is heated with caustic potash.
(*c*) Concentrated nitric acid acts on copper chips.
(*d*) Concentrated sulphuric acid is heated with nitre.

4. You are provided with five jars containing oxygen, hydrogen, sulphur dioxide, carbon dioxide and nitrogen. How would you identify which is which?

Or,

How is hydrochloric acid prepared? Describe briefly its properties and give its uses

METALS

CHAPTER XVII.

GENERAL PROPERTIES.

1. **Metals.**—Elements have from very early times been divided into metals and non-metals, the distinction being largely based on physical properties. The solids, which are hard, have a shining appearance, and can be easily beaten into sheets or drawn into wires, were termed **metals**; while all gases and light and brittle solids like carbon, sulphur and phosphorus were called **non-metals.** As some of the metals and non-metals were found to possess similar properties, this distinction could be no longer strictly observed. The present classification of elements into metals and non-metals is based on chemical as well as physical properties.

There can be no hard and fast line drawn between metals and non-metals, for the physical properties of metals vary between wide limits. The important properties of metals are:—

Physical.—Metals are all *opaque, e.g.* iron, lead, copper, etc. In thin films, however, some of them, such as gold leaves, are more or less translucent.

2. They have a peculiar lustre, known as *metallic lustre*. Thus, a piece of silver when freshly cut has a silvery-white lustre.

3. They have usually a high *specific gravity*. For instance, gold is about 19 times as heavy as water. Sodium and potassium, on the other hand,

which are decidedly metals, are lighter than water.

4. They are generally *hard, e.g.* iron, zinc and copper. Sodium and potassium are, however, so soft that they can be cut with a knife.

5. They are usually *malleable, i.e.* can be beaten into sheets, and *ductile, i.e.* can be drawn into wires. Gold and silver can be beaten into extremely thin sheets and drawn into very fine wires. Silver sheet ·00001 inch and wire ·0003 inch thick have been made.

6. They generally have a *high melting point.* Copper melts at 1083°C. and silver at 960°C. Mercury is the only metal liquid at the ordinary temperature.

7. They are usually *good conductors* of heat and electricity, *e.g.* copper and silver. Graphite, though a non-metal, is a good conductor of heat and electricity.

In each of the above properties we find that there are important exceptions, and we can hardly take any one property as a characteristic of metals.

Chemical.—1. Metals generally *dissolve in acids* like hydrochloric and sulphuric, giving off hydrogen gas. For example, zinc, magnesium and iron are readily acted upon by these acids with the evolution of hydrogen.

2. They form *oxides* which, when dissolved in water, have usually a basic reaction and form salts with acids. For instance, sodium, potassium and magnesium form basic oxides; while sulphur, phosphorus and carbon, which are non-metals, form acidic oxides (page 44). This is a very important distinction between metals and non-metals.

The general differences between metals and non-metals are summarised in the following table:—

No.	Metals.	Non-metals.
1.	Opaque. ...	Many (gases) are transparen , some (solids) are opaque.
2.	Metallic lustre. ...	No metallic lustre.
3.	High specific gravity	Low specific gravity.
4.	Solids with high melting point (except mercury).	Gases or solids with low melting point.
.	Generally malleable and ductile.	When soild, generally brittle.
.	Good conductors c heat and e ectricity.	Bad conductors of heat and electricity.
7.	Generally dissolve in acids, giving off hydrogen.	Do not easily dissolve in acids.
8.	Form basic oxides.	Form acidic oxides.

2. **Occurrence of metals.**—Only a few metals are found free in nature, *e.g.* copper, silver and gold. When a metal occurs free in nature, it is said to occur **native.** The solid elements and their compounds, which occur in the earth's crust, are known as **minerals.** Those minerals from which metals can be profitably extracted are called **ores** of the metals.

3. **Alloys.**—In many cases two or more metals, when melted together, dissolve in one another and form homogeneous, metallic-looking mixtures or compounds known as **alloys.** For example, copper and zinc form brass; copper, tin and zinc give **bronze.** The alloys are sometimes used in preference to metals, because they are generally harder and have a lower melting point than the pure metals. Their properties mostly depend upon the nature and the proportion of the metals present.

Alloys, in which mercury forms one of the constituents, are termed **amalgams,** *e.g.* **tin amalgam**

(tin and mercury) and sodium amalgam (sodium and mercury). Zinc plates are amalgamated by rubbing with mercury before they are used in batteries (see Part I, Chapter XLIII).

The following table gives the names, composition, and properties of some of the important alloys:—

Name of the alloy.	Component parts.	Properties.
Brass (ordinary).	Copper 2, zinc 1.	Harder than copper, casts well, and easily workable.
Dutch metal.	Copper 4, zinc 1.	Highly malleable.
Gun metal.	Copper 9, tin 1.	Very tough and tenacious.
Bell metal.	Copper 4, tin 1.	Hard, brittle and sonorous.
Bronze.	Copper 95, tin 4, zinc 1.	Fusible, casts well, does not corrode easily.
German silver.	Copper 3, zinc 2, nickel 1.	White in colour, tough and malleable.
Solder.	Lead 1, tin 1 (composition variable).	Soft and fusible.
Pewter.	Lead 1, tin 4.	Easily fusible, expands on cooling.
Type metal.	Lead 4, antimony 1.	Fusible, expands on cooling.
Gold coin (British).	Gold 11, copper 1.	Harder than gold.
Silver coin (Indian).	Silver 11, copper 1.	Harder than silver

QUESTIONS.

1. Into what two classes are elements divided? What are the characteristic properties of each class?

2. What are the principal characteristics of metals? Do you consider the division of elements into metals and non-metals satisfactory? [P. U. Matric., 1923.]

3. You are given an unknown element. How will you show whether it is a metal or a non-metal?

4. Classify the following elements into metals and non-

metals :—

Carbon, sodium, phosphorus, nitrogen, lead, chlorine and zinc.

5. What is meant by the terms: *ore, mineral* and *native state?*

6. What is an alloy? Name four important alloys and give their composition and uses.

7. Which metals are present in a (*i*) pice, (*ii*) rupee, and (*iii*) sovereign?

CHAPTER XVIII.

SODIUM AND POTASSIUM.

Sodium.

4. **Sodium** was discovered by Sir Humphrey Davy who obtained it in 1807 by the electrolysis of caustic soda.

5. **Occurrence.**—Sodium does not occur free in nature, as it readily combines with oxygen. In the combined state, it occurs mainly as (*i*) sodium chloride or common salt (*namak*) in sea-water, salt-lakes and salt mines, (*ii*) sodium nitrate or Chili saltpetre in Chili and Peru in South America, and (*iii*) sodium silicate in many minerals.

6. **Preparation.**—Sodium is prepared on a large scale by the electrolysis of fused caustic soda in a big cylindrical iron vessel. On passing the current through electrodes placed in the vessel, the caustic soda decomposes into sodium, oxygen and hydrogen. Sodium collects on the cathode, while oxygen and hydrogen escape—

$$\text{Caustic soda} = \underbrace{\text{sodium} + \text{hydrogen}}_{\text{cathode}} + \underbrace{\text{oxygen.}}_{\text{anode}}$$

7. **Properties.** *Physical.*—Sodium is a soft white metal which can be easily cut with a knife. Its specific gravity is ·98, and so it floats on water. When freshly cut, it has a white metallic lustre which gradually disappears on exposure.

Chemical.—Sodium is a very reactive metal having a great tendency to combine with oxygen.

Sir Humphrey Davy (1778—1829).

(*To face page* 162).

When exposed to air, it readily takes up oxygen and a brownish coating of sodium oxide is formed on its surface. On account of its strong affinity for oxygen, sodium is kept under kerosene oil which contains no oxygen. It reacts vigorously with water, forming caustic soda and evolving hydrogen gas (expt. 47, page 57)—

Sodium+water = sodium hydroxide+hydrogen.

Experiment 130.—Clamp a glass tube open at both ends in a vertical position with its lower end dipping about 3 cm. below the surface of water contained in a dish (Fig. 56). Add some litmus solution, which has been reddened by the addition of a drop of an acid, to the water in the dish. Drop a small piece of metallic sodium, carefully dried between the folds of a filter-paper, into the clamped glass tube. A vigorous action takes place accompanied by the evolution of hydrogen gas, which may be ignited at the upper end of the tube. The sodium hydroxide formed, being an alkali, changes the colour of the litmus solution from red to blue.

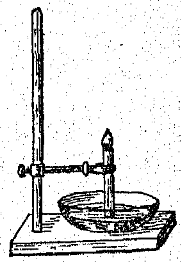


Fig. 56.—Action of sodium on water.

8. **Uses.**—Sodium is used to extract water from ether. Sodium amalgam (sodium and mercury) is used in preference to sodium in the preparation of hydrogen gas to make the action easily controllable (page 57).

Compounds of Sodium.

9. **Sodium hydroxide or caustic soda.**—As already mentioned, caustic soda is formed when sodium reacts with water. In India it is prepared by adding slaked lime to a boiling solution of

washing soda in an iron vessel —

Slaked lime +washing soda
(calcium hydroxide) (sodium carbonate)
= calcium carbonate+sodium hydroxide.

The insoluble calcium carbonate is removed by filtration, and the filtrate containing mostly caustic soda is evaporated to dryness in iron pans. The residue is fused and sometimes cast into sticks. The process, though cheap and simple, does not yield a pure product.

Sodium hydroxide is now largely prepared by the electrolysis of a solution of common salt (sodium chloride) in water. On passing the current through the solution, sodium and chlorine are liberated at the cathode and anode respectively. The metallic sodium then reacts with water to form caustic soda. Since caustic soda and chlorine mutually react when brought into contact, special devices are used to keep the two apart.

Caustic soda is a white amorphous solid which readily absorbs moisture and carbon dioxide from the air. Sodium hydroxide is known as *caustic soda* on account of its caustic or corrosive action on the skin. It dissolves readily in water with the evolution of heat, and the solution has a strong alkaline reaction towards litmus. Caustic soda is largely employed in the manufacture of soap and paper. It is also used as a reagent in the laboratory. A solution of caustic soda in water is employed for the detection of cotton in a woollen cloth.

Experiment 131.—Dissolve 2 grams of caustic soda in 100 grams of water. Put a piece of a woollen cloth suspected to contain cotton, and boil for about ten minutes. If the cloth dissolves completely, it is pure wool. If a

residue is left, the cloth is a mixture of cotton and wool.

10. **Sodium chloride or common salt** (*namak*).

The sources of supply of common salt may be divided into three main headings:—

1. Sea-water, which yields about 62 per cent. of the total production of the world.

2. Lake-water, which is an important source of supply in India. A large quantity of common salt used in India is obtained from the Sambhar salt-lake in Rajputana.

3. Rock-salt beds, such as the salt range of the Khewra mines (Punjab), the Kohat hills, the N. W. F. Provinces, and Mandi State in the Kangra district. This forms by far the most important source of common salt in India.

It also occurs abundantly in the salt deposits of Stassfurt (Prussia), Cheshire (California) and other places.

Common salt is obtained by different methods in different parts of the world:—

(1). In hot countries like India, the heat of the sun is used in the evaporation of water. When sea-water is exposed to the sun's heat and wind, the water is gradually evaporated and common salt remains behind as a white crystalline mass.

(2). In cold countries like England, where evaporation of water by exposure is not possible and evaporation by fuel is very expensive, most of the water is removed by freezing. On cooling the sea-water, pure water is gradually frozen out and the remaining solution becomes more and more concentrated. When the solution is sufficiently rich in common salt, it is evaporated over a fire.

(3) From rock-salt beds, common salt is ob-

tained by direct mining operations.

Sodium chloride forms colourless crystals which on heating break into smaller pieces with a crackling noise. It is an essential article of food for men and animals. Cows and other animals are often seen licking big pieces of common salt placed in streets. It is used in the preparation of almost all other sodium compounds, and is commonly employed in the manufacture of soap and hydrochloric acid. It acts as a preservative of butter, meat and other articles of food. In medicine it is occasionally used to cause vomiting.

11. **Sodium carbonate** or washing soda (*sajji*). The manufacture of sodium carbonate from common salt forms one of the most important industries of the day. The three methods commonly employed are: (*i*) Leblanc process, (*ii*) Solvay process, and (*iii*) electrolytic process:—

I. *Leblanc process.*—Common salt is first converted into sodium sulphate (*salt cake*) by heating it strongly with sulphuric acid—

Sodium chloride+sulphuric acid
=sodium sulphate+hydrochloric acid gas.

The hydrochloric acid gas evolved is passed through towers containing running water, where it is absorbed and forms hydrochloric acid.

The salt cake is then mixed with limestone and charcoal and heated in a large furnace. The resulting product consists mainly of calcium sulphide and sodium carbonate, and is known as *black ash*—

Sodium sulphate+calcium carbonate+carbon
=sodium carbonate + calcium sulphide
+ carbon dioxide.

The black ash is treated with warm water to

dissolve out the sodium carbonate. The solution is finally concentrated and allowed to crystallise. The calcium sulphide left behind forms an important source of sulphur.

This method is now obsolete and has been replaced by the other two methods.

II. *Solvay or ammonia-soda process.*—This process consists in passing carbonic acid gas under pressure into a strong solution of common salt saturated with ammonia, when sodium bicarbonate and ammonium chloride are formed—

Sodium chloride solution+ammonia+carbon dioxide
= sodium bicarbonate+ammonium chloride.

The sodium bicarbonate separates out as a white solid, leaving the more soluble ammonium chloride in solution. The bicarbonate is then removed and heated strongly to convert it into the carbonate—

Sodium bicarbonate
= sodium carbonate+water+carbon dioxide.

III. *Electrolytic process.*—Sodium chloride is first converted into caustic soda by electrolysis (page 164). A regular stream of carbon dioxide is then passed through the concentrated solution of caustic soda, when sodium carbonate is formed—

Sodium hydroxide+carbon dioxide
= sodium carbonate+water.

Properties.—Sodium carbonate is a white crystalline solid very soluble in water. It separates out from the aqueous solution in the form of crystals which contain nearly 63 per cent. of water. On exposure it *effloresces, i.e.* gradually loses water and falls to a powder. It is largely used under the name of **washing soda** for cleansing and washing purposes. It is also employed in the manufacture of soap.

12. **Sodium bicarbonate.**—As described above, this is obtained as a by-product in the Solvay process. It is also formed when carbon dioxide is passed into a strong solution of washing soda (sodium carbonate)—

Sodium carbonate+carbon dioxide+water
= sodium bicarbonate.

Sodium bicarbonate is a white powder much less soluble in water than the carbonate. When heated, it gives off carbon dioxide and is converted into the carbonate—

Sodium bicarbonate
= sodium carbonate+carbon dioxide+water.

It is therefore used in baking (*e.g.* in making *dabal roti*) under the name of *cooking soda* or *baking soda*. Like sodium carbonate, it is decomposed by an acid with the evolution of carbon dioxide. It is commonly used in medicine in cases of indigestion.

13. **Sodium nitrate** occurs abundantly in the desert regions of Chili and Peru, and is therefore known as *Chili saltpetre*. The crude nitrate is dissolved in water and the solution allowed to stand, when the insoluble earthy impurities settle down. The clear solution is then decanted off and allowed to crystallise.

Sodium nitrate is a white crystalline substance very soluble in water. It absorbs moisture from the atmosphere, and cannot therefore be used in place of potassium nitrate in the manufacture of gunpowder. It is largely used as a manure and in the manufacture of nitric acid.

14. **Sodium sulphate,** or Glauber's salt (*khari*). This compound occurs abundantly in Behar and the United Provinces, and is known as *Patna khari*.

It is obtained as a by-product in the manufacture of sodium carbonate by the Leblanc process (page 166). Like sodium carbonate, its crystals effloresce on exposure to air.

Sodium sulphate is largely used in the manufacture of sodium carbonate and glass. The pure salt is used in medicine.

Potassium.

15. **Potassium.**—Like sodium, potassium does not occur free in nature. It is largely found in the combined state as: (*i*) potassium chloride (*sylvite*) and a double chloride of potassium and magnesium (*carnallite*), (*ii*) potassium nitrate (nitre, *shora*) in Bengal saltpetre. Wood ashes contain about 30 per cent. of potassium carbonate.

The metals sodium and potassium and their compounds resemble one another so closely in their methods of preparation and properties that it is hardly necessary to devote a separate chapter to the discussion of potassium and its compounds. Like sodium, the metal potassium is prepared by the electrolysis of fused caustic potash—

$$\text{Caustic potash} = \underbrace{\text{potassium} + \text{hydrogen}}_{\text{cathode}} + \underbrace{\text{oxygen}}_{\text{anode}}$$

Potassium collects on the cathode, while the hydrogen ecsapes.

Like sodium, potassium is a soft white metal lighter than water. It has a strong affinity for oxygen, and like sodium it is also kept under kerosene oil. It decomposes water so vigorously that the evolved hydrogen catches fire.

16. **Caustic potash,** like caustic soda, is a white

deliquescent solid having a corrosive action on the skin. It readily dissolves in water and forms a strongly alkaline solution. It is largely used in the manufacture of toilet soap.

17. **Potassium carbonate** (*Jau khar*).—This compound has long been obtained exclusively from the wood ashes. The wood is burnt in pits, the ashes treated with water, and the solution evaporated to dryness. The crude salt thus obtained is known as *potash*, and the pure product as *pearlash*.

It is now largely manufactured by the Leblanc process. The Solvay process cannot be employed, *as potassium bicarbonate is very soluble in water.*

Potassium carbonate crystallises out from the aqueous solution in the form of long prisms. It is highly deliquescent, and like sodium carbonate its solution is alkaline towards litmus. It is used in making soft soap and hard glass. In India crude potash is largely used as a fertilizer, especially for rice crops.

18. **Potassium nitrate** or nitre (*shora*).—This occurs as a white crust on the soil in hot countries like India. It is found in large quantities in the Punjab, Bengal and the United Provinces. It is obtained from the nitrous earth (*shorili matti*) which is dissolved in water and the solution passed through filters of porous earth, grass, etc. The clear solution is then concentrated in iron pans (*karahis*) and allowed to crystallise.

Potassium nitrate is prepared from the cheaper sodium nitrate by treating the latter with potassium chloride—

Sodium nitrate+potassium chloride
= sodium chloride+potassium nitrate.

On concentrating the solution the less soluble

sodium chloride separates out, whereas potassium nitrate remains behind. The solution is filtered while hot to remove the sodium chloride and the filtrate allowed to cool, when crystals of potassium nitrate are deposited.

Nitre is a white crystalline solid which, when dissolved in water, causes a fall of temperature. When heated strongly, it decomposes to give oxygen and is therefore used in the manufacture of gunpowder, matches and fireworks. Gunpowder consists of a mixture of nitre, charcoal and sulphur. On firing the powder, carbon and sulphur burn with an explosive noise, the necessary oxygen being supplied by the nitre.

Experiment 132.—Powder some potassium nitrate, charcoal and sulphur in *separate* mortars. Mix together 15 parts of the nitrate, 3 parts of charcoal and 2 parts of sulphur by a feather on a piece of paper. Place the mixture on a stone slab and apply a flame to it. The mixture burns brightly.

19. **Potassium chlorate** was formerly obtained by passing chlorine into a hot solution of caustic potash. The old process has now been replaced by the electrolytic method. A hot solution of potassium chloride is electrolysed in a large cast iron vessel, and the temperature of the solution kept at about 50°C. Potassium chloride at first decomposes into potassium and chlorine. The liberated potassium reacts with water, forming potassium hydroxide which further reacts with chlorine, forming potassium chlorate—

Potassium chloride = potassium + chlorine,

Potassium + water

= potassium hydroxide + hydrogen,

Potassium hydroxide + chlorine
= potassium chloride + potassium chlorate + water.
The chlorate crystallises out first when the solution cools and is thus separated from potassium chloride.

Potassium chlorate is a white crystalline solid. When heated, it decomposes to give out oxygen—

Potassium chlorate
= potassium chloride + oxygen.

Hence it is largely used in the preparation of oxygen (page 39), and in the manufacture of matches and fireworks. It is also used in medicine, especially in cases of throat trouble.

20. **Potassium permanganate** is a dark-violet, crystalline solid containing potassium, manganese and oxygen. It is a powerful disinfectant and antiseptic, and is frequently dropped in wells and water-tanks to kill the disease germs present in the water. *Candy's fluid* (*lal dawai*) containing 1 part of potassium permanganate in 50 parts of water is a purple-coloured liquid which is largely used as a disinfectant. It is also employed as a valuable laboratory reagent.

QUESTIONS.

1. Give a brief account of the preparation, properties and uses of sodium.
2. Describe the action of sodium and potassium on water.
3. How does caustic soda differ from washing soda? Give the uses of the two compounds
4. Describe the important properties of caustic potash. What happens to it on exposure to air? What is the difference in the action of hydrochloric acid on fresh caustic potash, and on the same after a long exposure to air?
5. Describe a cheap method of preparing caustic soda in large quantities, and explain the chemical changes which occur in this process. State the important properties and uses of this

substance. [P. U. Matric., 1918.]

6. Name the important sources of common salt. How is it obtained from each of these sources?

7. Starting with sodium chloride, how would you prepare (*i*) washing soda, (*ii*) sodium sulphate, and (*iii*) chlorine?

8. What is *baking soda?* How can it be converted into washing soda? What are its uses?

9. How would you separate the constituents of gunpowder and identify them?

CHAPTER XIX.

COPPER AND SILVER.

Copper (*tamba*).

21. **Copper** was one of the earliest metals known to the ancients, for it occurs free in nature. It was formerly obtained in considerable quantities in Southern India and Rajputana, and was largely used for the manufacture of utensils, war weapons and coinage.

22. **Occurrence.**—Copper is found native in Lake Superior, Japan, Austria and India. The most valuable ores are:—

No.	Common name.	Chemical name.	Constituents.
1	Copper pyrites.	Copper and iron sulphide.	Copper, iron and sulphur.
2	Cuprite or ruby copper.	Copper oxide.	Copper and oxygen.
3	Copper glance.	Copper sulphide.	Copper and sulphur.

23. **Properties.** *Physical.*—Copper is a bright-red metal having a specific gravity of 8·9. It is fairly hard, and can be beaten into thin sheets and drawn into fine wires. Metals which, like copper, can be drawn into wires are called *ductile,* while those which, like gold and silver, can be beaten into leaves, are termed *malleable.* Next to silver, it is the best conductor of heat and electricity.

Chemical.—In dry air copper does not tarnish,

but in the presence of moisture and carbon dioxide it becomes coated with a greenish layer of the carbonate, commonly known as verdigris (*zangar*). When heated in air, it changes into black oxide of copper.

Dilute hydrochloric acid has no action on the metal. Even hot and concentrated acid does not attack copper in the absence of air. But when air is present, the metal is slowly dissolved by the acid, forming copper chloride—

Copper+hydrochloric acid+oxygen (of air)
= copper chloride+water.

Sulphuric acid, dilute or strong, has no action on copper in the cold. But the hot and concentrated acid dissolves the metal with the formation of copper sulphate and sulphur dioxide—

Copper+sulphuric acid
= copper sulphate+sulphur dioxide+water.

Nitric acid acts readily even in the cold with the formation of copper nitrate and reddish-brown fumes of nitrogen peroxide—

Copper+nitric acid
= copper nitrate+nitrogen peroxide+water.

24. **Uses.**—Next to iron, copper is the most useful metal. On account of its resistance to the action of air, it is extensively used in making household utensils, boilers, stills and coins. The high conductivity of copper is made use of in making electric apparatus and telegraph wires.

25. **Alloys.**—Copper forms a number of useful alloys: (1) **brass** (*pital*) contains copper and zinc and is used in making utensils, tools and idols. The huge brass idols of Burma and the ceremonial

tools and vessels manufactured in several parts of India are quite familiar; (2) **bronze** (*kansi*) contains copper, zinc and tin, and is used in making utensils, coins and statues; (3) **German silver,** containing copper, zinc and nickel, is used for making utensils and imitating silver.

26. **Copper sulphate.**—The most important compound of copper is copper sulphate or blue vitriol (*nila tutia*). It is formed by heating copper with strong sulphuric acid. It crystallises out from the aqueous solution in blue crystals containing about 36 per cent. of water. On heating, the crystals lose this water and fall to a white amorphous powder. This water, which is always present in the crystals of copper sulphate, is termed **water of crystallisation.** Copper sulphate crystallises out with five molecules of water of crystallisation. Many other salts, such as iron sulphate, zinc sulphate, sodium carbonate, etc., crystallise out from the aqueous solution with a definite number of molecules of water of crystallisation. The white powder, anhydrous copper sulphate, obtained by heating copper sulphate crystals, regains its colour on the addition of water.

Experiment 133.—Take a few crystals of copper sulphate and powder them. Heat the powder in a test-tube. Water is given off in the form of vapour which condenses on the upper part of the tube, and the copper sulphate gradually turns white. Allow the tube to cool and add a few drops of water to the white powder; the blue colour reappears.

Copper sulphate is used in batteries (Daniell's cell), copper plating, and in the preparation of almost all other compounds of copper. In medicine, it is used as a caustic in granular lids (*kukre*)

and in the manufacture of insecticides. Mixed with milk of lime, it is used as a spray for killing moulds and scales on fruit trees.

Silver (*chandi*).

27. On account of its attractive appearance and occurrence in the free state, silver has been known from times immemorial.

28. **Occurrence.**—Silver is found free in nature mixed with small quantities of copper and gold. It is also present in lead ores, which form an important source of the metal. In the combined form it occurs as silver sulphide and chloride—

Common name.	Chemica name.	Constituents.
Silver glance.	Silver sulphide.	Silver and sulphur.
Horn silver.	Silver chloride.	Silver and chlorine.

29. **Properties.**—Silver is a lustrous white metal, very malleable and ductile. It is the best conductor of heat and electricity, but it is too costly to be used for such purposes. It is about 11 times as heavy as water. It is unaffected by air but forms black silver sulphide with sulphuretted hydrogen, a compound of hydrogen and sulphur. Silver ornaments generally turn black owing to the action of sulphuretted hydrogen present in air. The black coating consists of silver sulphide. Silver is not acted upon by hydrochloric or sulphuric acid, but it dissolves in nitric acid with the formation of silver nitrate—

Silver+nitric acid

= silver nitrate+water+nitrogen peroxide.

30. **Uses.**—On account of its beautiful appearance and resistance to the action of air, silver is used for making vessels, ornaments and mirrors. Alloyed with copper, it becomes hard and is used for coins. Silver is also used for coating various articles made of cheaper metals. The process by which this is done is known as *electroplating* (see Part I, Chapter XLVI).

31. **Silver nitrate.**—This is an important compound of silver, and is obtained, as mentioned above, by the action of nitric acid on silver. It is a colourless crystalline substance soluble in water. The solution turns black on exposure to light, and leaves a black stain when dropped on the skin or cloth. Silver nitrate is used in the preparation of marking ink and hair-dyes. In medicine, its solution is used under the name, *caustic lotion*, as a cure for granular lids. In the laboratory, silver nitrate solution is generally used for testing hydrochloric acid or a chloride.

Silver salts are very sensitive to light and are used in photography.

QUESTIONS.

1. Name the important ores of copper and give their chemical composition.
2. Describe the properties of copper. What is the action of common acids on copper?
3. Give the composition, properties and uses of alloys of copper.
4. What is 'verdigris'? How is it formed? Why is it occasionally seen on copper coins?
5. Explain the changes that take place when (*i*) blue vitriol is heated in air, and (*ii*) when water is added to the heated mass.
6. Describe the properties and uses of silver.

7. Why do silver ornaments turn black when exposed to air for a long time?

8. State the composition, important properties and uses of brass, silica and marble. [P. U. Matric, 1919.]

9. Write what you know about the composition, properties and uses of the following:—(*a*) marble, (*b*) quartz, (*c*) Chili saltpetre, and (d) German silver. [P. U. Matric, 1924.]

CHAPTER XX.

ZINC AND MERCURY.

Zinc (*jast*).

32. **Occurrence.**—Zinc does not occur in the free state in nature. The chief ores are:—

No.	Common name.	Chemical name.	Constituents.
1.	Zinc blende	Zinc sulphide	Zinc and sulphur.
2.	Calamine	Zinc carbonate	Zinc, carbon and oxygen.
3.	Zincite	Zinc oxide	Zinc and oxygen.

33. **Properties.**—Zinc is a bluish-white, hard and crystalline metal. At the ordinary temperature it is brittle, but at 100° to 150°C. it becomes soft, malleable and ductile. When heated in air, it burns with a bluish flame, forming zinc oxide. Zinc dissolves readily in dilute hydrochloric and sulphuric acids, giving hydrogen (expt. 49).

34. **Uses.**—On account of its lightness and resistance to the action of air and moisture, *sheet zinc* is largely used for roofing and for making cisterns, baths, etc. *Galvanised iron,* made by dipping iron sheets in molten zinc, is commonly used for roofs tanks, pipes, etc. Zinc is also used in galvanic batteries and in the formation of alloys.

35. Alloys.—Zinc forms several alloys:—(*i*) **Brass** consisting of copper and zinc, (*ii*) **German silver** of copper, zinc and nickel, and (*iii*) **Dutch metal** of copper and zinc.

36. **Compounds of zinc.**—The two important

compounds of zinc are: zinc oxide and zinc sulphate.

1. **Zinc oxide** is commonly prepared in villages by burning zinc in clay crucibles. The burnt mass is finely powdered and used as a cure for sore eyes. Zinc oxide is also used in paints and medicine in the form of zinc ointment.

2. **Zinc sulphate** or white vitriol is a white crystalline solid, obtained by the action of sulphuric acid on zinc. It is soluble in water, and the dilute solution is used under the name *zinc lotion* as a cure for sore eyes.

Mercury (*para*).

37. **History.**—The element has long been known as quicksilver or liquid silver (*hydrargyrum*), as it looks like silver and flows like water. It was supposed by ancient Greeks and Arabians to be a constituent of all metals and to give them the property of volatility.

38. **Occurrence.**—Mercury occurs in small quantities in the free state, but is usually found as *cinnabar* (mercury sulphide).

39. **Purification of mercury.**—Mercury, as generally found in the market, is impure and contains traces of lead, zinc, copper and other impurities. It may be cleaned by

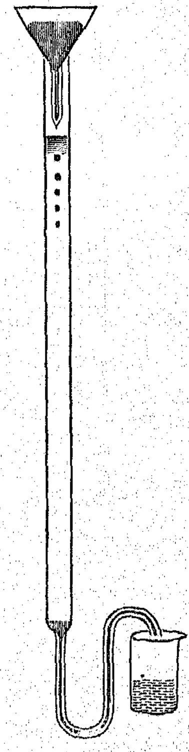


Fig. 57.—Purification of mercury.

filtration through chamois leather, and then purified by running a fine stream of the filtered metal down a long column of dilute nitric acid (Fig. 57). The nitric acid almost completely dissolves the impurities, and sufficiently pure mercury is obtained. The metal is then washed with water and dried in a hot-air oven. If a high degree of purity is required, it is distilled in vacuum.

40. **Properties.** *Physical.*—Mercury is a silver-white metal having a specific gravity of 13·59. It is the only metal liquid at the ordinary temperature. It solidifies at $-39°$C. and boils at $357°$C. Impure mercury leaves a metallic streak '*tail*' behind when rolled over a glass plate.

Chemical.—The metal is not affected by air at the ordinary temperature, but when heated to about $300°$C. it forms red oxide of mercury. On further heating, the oxide decomposes to give back mercury and oxygen. It is not acted upon by hydrochloric or dilute sulphuric acid. Concentrated sulphuric acid and nitric acid, however, dissolve the metal on heating, forming the corresponding salts—

Mercury+sulphuric acid
= mercury sulphate+sulphur dioxide+water.

Mercury+nitric acid
= mercury nitrate+nitrogen peroxide+water.

41. **Amalgams.**—Mercury has the property of dissolving many metals and forming *amalgams*. Sodium amalgam is used in the preparation of hydrogen (page 57). Tin amalgams is used in the formation of mirrors. Amalgams of gold, copper and zinc are used in dentistry for stopping teeth.

42. **Uses.**—On account of its high specific gravity and low vapour pressure, mercury is ex-

tensively used in barometers. Its uniform expansion favours its use in thermometers (see Part I, Thermometry). Its tendency to form amalgams is utilised in the extraction of metals like gold and silver from their ores.

43. **Mercury oxide.**—This is obtained by heating mercury in air—

Mercury + oxygen = mercury oxide.

It may also be obtained by strongly heating nitrate of mercury—

Mercury nitrate
= mercury oxide+nitrogen peroxide+oxygen.

Mercury oxide is used in red ointment to cure certain ulcers.

QUESTIONS.

1. Name the principal ores of zinc and mercury. State the chemical name and composition of each.
2. What are the chief physical and chemical properties of zinc and mercury?
3. What are the important uses of mercury? Why is mercury used in barometers and thermometers in preference to other liquids?
4. State the difference between *amalgams* and *alloys*. Name the important alloys of zinc. State their composition and uses.
5. What is *galvanised* iron and why is iron galvanised?

CHAPTER XXI.

ALUMINIUM.

44. **Occurrence.**—Aluminium does not occur free in nature, but in the combined state it is very widely distributed. It occurs mostly as oxide in three different forms: (*i*) *corundum* in Mysore and Madras, (*ii*) *ruby* in Burma, and (*iii*) *sapphire* in Kashmir. *Bauxite* (aluminium oxide) and *cryolite* (fluoride of sodium and aluminium) form the chief sources of the metal. Bauxite of excellent quality exists in western Chota Nagpur, though it is found there in small quantities. The metal also occurs as a silicate in *clays, felspar, granite* (*sang. khara*) and *mica* (*abraq*).

45. **Properties.** *Physical.*—Aluminium is a bluish-white metal having a specific gravity of 2·6. It is very malleable and ductile but requires frequent annealing (*abdari*) during the process of hammering or drawing. It is as good a conductor of heat and electricity as copper, weight for weight.

Chemical.—The metal remains almost unchanged in dry air, but in moist air a superficial layer of the oxide is formed and this protects the metal from further action. It does not decompose water even at its boiling point. Nitric acid has almost no action on the metal, but dilute hydrochloric acid dissolves it rapidly. Concentrated sulphuric acid attacks the metal with the evolution of sulphur dioxide—

Aluminium+sulphuric acid

= aluminium sulphate+sulphur dioxide+water.

It is readily dissolved by caustic soda or potash, and it is for this reason that aluminium utensils can be easily cleaned by wood ashes which contain potassium carbonate, a weak alkali.

On account of its strong affinity for oxygen aluminium forms a powerful reducing agent. Oxides of certain metals, such as chromium or manganese, which cannot be reduced by charcoal, are easily reduced to the metallic state when heated with aluminium powder.

Experiment 134.—Take a mixture of aluminium powder and black copper oxide on an iron spatula, and hold it in a Bunsen flame. The action takes place immediately and is accompanied by instantaneous flashes of light.

This mixture of aluminium powder and copper oxide is coated on a thin wire of copper or iron and is employed in making *phuljaris*.

If aluminium powder be mixed with sodium peroxide and the mixture treated with water, the oxygen produced by the action of water on the peroxide is easily taken up by aluminium and a violent action takes place.

Experiment 135.—Place some sodium peroxide in a clean and dry test-tube, and add an equal quantity of aluminium powder. Mix the two powders well by shaking and clamp the test-tube behind a screen. Allow a drop or two of water to flow from a long glass tube into the test-tube. A violent action takes place with explosion that shatters the test-tube.

46. **Uses.**—Aluminium and its alloys are used for making photographic cameras, household utensils, telephone receivers and science apparatus. On account of its lightness and attractive appearance it it used for making trays, picture frames, combs, and for many other ornamental purposes.

Aluminium powder, when mixed with linseed oil, forms a silvery paint, *aluminium paint,* which is sometimes used for covering electric and telegraphic poles. It is also employed as a strong reducing agent.

47. **Alloys.**—Aluminium forms several valuable alloys which have a high tensile strength, and are less easily corroded than the pure metal. The most important alloys are: (*i*) **aluminium bronze** containing aluminium and copper, and (*ii*) **magnalium** containing aluminium and magnesium. Magnalium is generally used for making cheap balances.

QUESTIONS.

1. Mention the chief sources from which aluminium can be obtained, and also name the naturally occurring oxides of aluminium.

2. Describe the chief physical and chemical properties of aluminium.

3. Name the principal alloys of aluminium and give the important uses of aluminium and its alloys.

CHAPTER XXII.

TIN AND LEAD.

Tin (*qal'i*).

48. **Occurrence.**—Tin rarely occurs in the free state. The most important ore from which nearly the whole of the metal is obtained is *tinstone* (tin oxide) which is mostly found in Cornwall, Saxony and Southern Burma.

49. **Properties.** *Physical.*—Tin is a silver-white metal of specific gravity 7·3 and melting point 232·7°C. It is harder than lead but softer than zinc. At the ordinary temperature it is malleable and ductile, but at 200°C. it becomes brittle. 'Grain tin' or 'granulated tin,' so common in the market, is obtained by allowing the molten tin to drop from a height.

Chemical.—Tin is not affected by air at the ordinary temperature, but if heated strongly it burns with an intense white light, forming tin oxide. Dilute hydrochloric acid has a slow action on tin, but the hot and concentrated acid dissolves the metal more readily and forms tin chloride—

Tin+hydrochloric acid
= tin chloride+hydrogen.

Sulphuric acid has no action in the cold, but the strong acid dissolves tin on heating—

Tin+sulphuric acid
= tin sulphate+water+sulphur dioxide.

Cold and dilute nitric acid gives tin nitrate and

ammonium nitrate, while the concentrated acid gives metastannic acid and nitrogen peroxide—

Tin+nitric acid

= metastannic acid+nitrogen peroxide+water.

50. **Uses.**—Since tin is unaffected by air at the ordinary temperature, it is used for protecting other metals, as in tinware vessels. Tin-plate or 'tin' is made by dipping clean iron into molten tin. Pure tinfoil is used for packing food-stuffs, oils and other articles. Common pins are brass wires coated with tin by shaking them in a solution of a tin salt.

Tin is also commonly employed in 'tinning' kitchen utensils. The vessel to be tinned is heated and rubbed with ammonium chloride, which readily decomposes into ammonia and hydrochloric acid gases. The acid dissolves the oxide which may be present as a thin layer on the surface of the vessel, and thus makes the surface quite clean. Rosin is then sprinkled, which melts and protects the surface from oxidation. Molten tin is then spread over the surface of the vessel with a coarse piece of cloth.

Alloys.—Tin forms a number of alloys. The most important of them are:—

Alloy.	Constituents.	Alloy.	Constituents.
1. Gun metal	Copper and tin.	3. Bronze	Copper, tin and zinc.
2. Bell metal	Copper and tin.	4. Solder	Tin and lead.

Tin amalgam is used in coating mirrors. The salts of tin are used in dyeing and calico-printing.

Lead (*sisa*).

51. **Occurrence.**—Lead is occasionally found free in small quantities. The important lead ores are:—

No.	Common name.	Chemical name.	Constituents.
1.	Galena	Lead sulphide	Lead and sulphur.
2.	White lead	Lead carbonate	Lead, carbon and oxygen.
3.	Anglesite	Lead sulphate	Lead, sulphur and oxygen.

Galena is the most abundant ore of lead and occurs in England and the United States. It forms the chief source of commercial lead. It must not be confused with black-lead (page 75) which is a form of carbon and does not contain any lead at all.

52. **Properties.** *Physical.*—Lead is a bluish-grey metal of specific gravity 11·4 and melting point 326°C. It is so soft that it can be cut with a knife and leaves a mark on paper when drawn across it. When scraped or freshly cut, it has a metallic lustre which soon disappears due to the formation of a thin layer of its oxide. It cannot be hammered into foil or drawn into fine wire, but it can be rolled into sheet.

Chemical.—When heated in air, lead first forms the yellow oxide or litharge (*murda sang*, expt. 6), and if the heating be continued, it takes up more oxygen from the air and forms red lead (*sandhur*). Mixed with white lead and linseed oil, red lead is used as a cement for joining gas pipes. Litharge and red lead, when mixed with charcoal and heated strongly, lose their oxygen and give back the metal.

Experiment 136.—Mix some red lead with powdered charcoal and heat the mixture in a crucible (Fig. 58). The carbon takes up the oxygen of the red lead, and metallic lead appears in the form of shining beads. Take out one bead and draw it across a paper; a mark is left on the paper. Take another bead and press it on the table with a pestle. It is flattened out into a cake, showing that the metal is malleable.

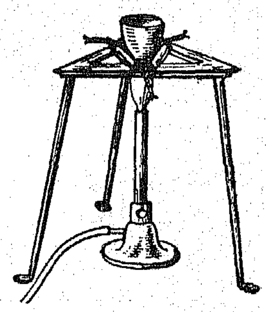


Fig. 58.—Reduction of red lead.

Hydrochloric, dilute or strong, and dilute sulphuric acid have no action on lead. But hot and concentrated sulphuric acid dissolves it with the formation of lead sulphate and sulphur dioxide—

Lead+sulphuric acid
= lead sulphate+water+sulphur dioxide.

Nitric acid attacks the metal, forming lead nitrate and nitrogen peroxide—

Lead+nitric acid
= lead nitrate+nitrogen peroxide+water.

Acids present in vegetables form soluble poisonous compounds with lead. Hence cheap tin-plated vessels, which sometimes contain lead, should never be used for cooking. Lead is displaced from its salts in solution by metallic zinc.

Experiment 137.—Suspend a piece of zinc in a solution of lead acetate or *sugar of lead* contained in a cylinder (Fig. 59). Allow the whole to stand for a few hours. Lead will be deposited on the zinc piece in the form of a beautiful tree, lead tree—

Lead acetate+zinc=zinc acetate+lead.

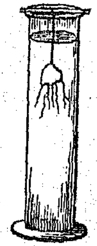


Fig. 59.—The lead tree.

53. **Uses.**—Lead is largely used in making pipes, for it can be made into long pieces which can be easily bent, cut and joined. On account of its very slow action with air, water and acids, *sheet lead* is used to cover roofs, to line sinks and cisterns, and to form chambers in the manufacture of sulphuric acid. About one-third of the total production of lead is employed in the manufacture of white-lead used for making paints.

54. **Alloys.**—Lead forms an important constituent of several alloys. **Solder** contains lead and tin, and is used for soldering. **Type metal** contains lead and antimony, and is used for making types. When alloyed with a little arsenic, lead becomes very hard and is used in making shot and bullets.

55. **Compounds of lead.**—All compounds of lead are poisonous. **Lead carbonate** or white lead (*sufeda*) is extensively used in making paints. **Litharge** (*murda sang*) is used in making paints and preparing other compounds of lead. **Red lead** (*sandhur*) is employed as a pigment and a cement (putty or *potin*, a mixture of red lead and white lead) for gas-pipes and steam joints. It is also used in the manufacture of matches and the construction of lead accumulators.

QUESTIONS.

1. Describe the important properties of tin. What are its chief uses?

2. Name the important ores of lead and give their chemical composition.

3. Describe the properties and uses of lead.

4. Give the composition and uses of some important alloys of tin and lead.

5. How are the following obtained:—(*i*) litharge from lead, (*ii*) red lead from litharge, and (*iii*) lead from red lead?

6. How do black lead, white lead, and red lead differ from one another?

7. Explain the chemical actions which take place when—

(1) hydrochloric acid acts on zinc,
(2) nitric acid acts on copper,
(3) sulphuric acid acts on common salt, and
(4) sodium acts on water. [P. U. Matric., 1916.]

CHAPTER XXIII.

IRON (*loha*).

56. **Iron** seems to have been known from very early times. The relics of early iron works, such as *lohe ki lat* near Delhi, that exist up to the present day, show that the manufacture of iron was practised by the ancient Hindus in no crude a fashion. Iron is now manufactured on a very large scale in several parts of the world. Tata Iron Works at Tatanagar near Calcutta alone produce several thousand tons of cast iron annually.

57. **Occurrence.**—Iron is one of the most widely distributed elements in nature. It occurs free in small quantities, but in the combined form it is found very abundantly in nature. The chief ores are:—

No.	Common name.	Chemical name.	Constituents.
1.	Magnetite or magnetic ore (*chamaq pathar*).	Iron oxide.	Iron and oxygen.
2.	Hæmatite.	Iron oxide.	Iron and oxygen.
3.	Clay iron stone.	Iron carbonate.	Iron, carbon and oxygen.
4.	Iron pyrites.	Iron sulphide.	Iron and sulphur.

58. **Varieties of iron.**—The commercial iron is usually mixed up with other substances, especially carbon, which greatly modify its properties. The three important varieties of iron are: cast iron, wrought iron and steel. Each variety has its own characteristic properties, depending upon the amount of carbon which it contains.

1. **Cast iron** contains 1·5 to 4·5 per cent. of carbon, and is hard and brittle. It cannot be *welded*, *i.e.* united by hammering red-hot pieces of iron, but it can be moulded into any shape. It is largely used for making castings, stoves, pipes, pillars, parts of machinery, and various other objects.

2. **Wrought iron.**—This is the purest form of iron, and contains less than 0·5 per cent. of carbon. It is tough and malleable. Unlike cast iron, it can be easily welded, rolled into sheets, and drawn into wires. It is used for making rods, nails, spikes, horse-shoes and other articles.

3. **Steel** (*faulad*).—It contains 0·5 to 1·5 per cent. of carbon, and its properties vary with the amount of carbon present. If the amount of carbon be small, steel is tough and ductile like wrought iron. But if it reaches 1·4 per cent., it is hard and brittle like cast iron. The properties of steel are, therefore, intermediate between those of cast and wrought iron. It is both fusible and malleable, and can be cast and welded. It is harder, stronger and more durable than wrought iron.

The properties of steel depend not only upon the proportion of carbon which it contains but also upon its final treatment. Thus, steel containing only one per cent. carbon, when heated to a high temperature and then suddenly cooled by immersing in cold water, becomes so hard that it can scratch glass, and so brittle that it can be easily broken into pieces. This process is known as **hardening**. When *hardened* steel is again heated to a very high temperature and then cooled slowly, it becomes soft, ductile and elastic, and the process is termed annealing (*abdari*).

Between these two limits, various degrees of hardness and elasticity can be obtained by heating steel to a high temperature and then cooling it at different rates. This process is known as **tempering.**

Steel cannot be magnetised so easily as cast or wrought iron, but when once magnetised, it retains its magnetism for a long time. It is for this reason used in making permanent magnets. It is also employed in making springs, tools, knives, razors, swords, etc.

59. **Pure iron.**—Chemically pure iron is unknown in commerce. It can be prepared in the laboratory by passing hydrogen gas over strongly heated iron oxide—

Iron oxide+hydrogen = iron+water.

Pure iron is a lustrous, white metal having a specific gravity of about 8. It is the most magnetic of all the metals. It remains unchanged in dry air, but in moist air it soon becomes covered with a reddish-brown film—*iron rust.* In order to prevent iron from rusting it is—

(*i*) Coated with tin or zinc. These metals are not easily affected by air. *Tinned iron* is prepared by dipping a clean sheet of iron in molten tin. It is mostly used in making kerosene oil tins.

Galvanised iron is made by dipping sheets of iron in molten zinc (page 179).

(*ii*) Covered with finely powdered graphite, or coated with vaseline or a vegetable oil. Knives, forks, swords and other valuable articles made of iron are generally coated with vaseline to prevent rusting in rainy weather.

(*iii*) Coated with enamel, nickel, or a paint. Bicycle handles, bells and pumps are generally

nickelled, whereas mud-guards and other parts are covered with a black paint. Enamelled plates, jugs, dishes and other utensils are frequently used for domestic purposes and do not rust unless the protective layer of the enamel has been removed.

Iron is readily acted upon by dilute hydrochloric and sulphuric acids with the evolution of hydrogen gas. It readily combines with sulphur, forming iron sulphide, and burns brilliantly in oxygen, forming iron oxide (expt. 40).

60. **Iron sulphate** or green vitriol (*hira kasis*) is obtained by the action of dilute sulphuric acid on iron. It is an important compound of iron, and is largely employed in the manufacture of inks and paints.

QUESTIONS.

1. Name the important ores of iron and give their chemical composition.

2. What are the distinguishing characters of cast iron, wrought iron and steel?

3. What is meant by the terms: *tempering, annealing* and *hardening?*

4. What is *tinned* iron and how does it differ from galvanised iron? Why is iron tinned?

5. What is the action of (*i*) air, (*ii*) steam, and (*iii*) sulphuric acid on iron?

6. Pieces of copper, zinc and iron are separately heated in a crucible over a burner. Describe what happens in each case.

7. State the properties of the following substances, and name the elements of which each is composed:—(*i*) black-lead, (*ii*) gypsum, (*iii*) Glauber's salt, (*iv*) steel, and (*v*) ozone. [P. U. Matric., 1918.]

8. Write what you know about the composition, properties and uses of the following: (*a*) washing soda, (*b*) graphite, (*c*) blue vitriol, and (*d*) quartz.

CHAPTER XXIV.

ACTION OF ACIDS ON METALS.

61. The action of an acid on a metal varies with the nature and concentration of the acid, and with the temperature. In this chapter, the action of the three most important acids, *viz.* hydrochloric, sulphuric and nitric, on some of the important metals will be briefly discussed.

62. **Hydrochloric acid.**—Sodium, potassium, aluminium and iron are easily dissolved by dilute hydrochloric acid at the ordinary temperature. The corresponding chloride is formed and hydrogen gas evolved, *e.g.*—

Iron+hydrochloric acid
= iron chloride+hydrogen.

Pure zinc is soluble in dilute hydrochloric acid only with difficulty, though the commercial zinc is readily dissolved by it. The concentrated acid reacts more vigorously with these metals and gives the same products.

Silver, mercury and lead are practically unattacked by hydrochloric acid, whether dilute or concentrated. Copper, however, is slowly dissolved by the concentrated acid in the presence of air, forming copper chloride.

63. **Sulphuric acid.**—Dilute sulphuric acid acts on sodium, potassium, zinc (commercial) and iron, forming the corresponding sulphate and giving off hydrogen, *e.g.*—

Iron+sulphuric acid = iron sulphate+hydrogen.

As in the case of dilute-hydrochloric acid, *pure* zinc is acted upon by dilute sulphuric acid only with difficulty. Copper, silver and mercury are not acted upon by the dilute acid.

Concentrated sulphuric acid reacts with all the metals that are acted upon by the dilute acid and also with several other metals which are not acted upon by the dilute acid. The action in this case takes place on heating and sulphur dioxide, but no hydrogen is evolved. Copper, silver, mercury, tin and lead all behave in this way and when heated with strong sulphuric acid, they give sulphur dioxide and form the corresponding sulphate at the same time—

Copper+sulphuric acid
= copper sulphate+sulphur dioxide+water.

64. **Nitric acid.**—This acid readily acts on most of the metals, forming nitrates. Unlike other acids, hydrogen is seldom evolved. As a matter of fact, hydrogen evolved is immediately oxidised by nitric acid which is itself reduced to oxides of nitrogen, *e.g.*—

Copper+nitric acid
= copper nitrate+hydrogen,

Hydrogen+nitric acid
= water+nitrogen peroxide.

Hence the final product is nitrogen peroxide and not hydrogen.

65. **Summary.**—The table given on the next two pages summarises the action of hydrochloric, sulphuric and nitric acids on some important metals.

No.	Metal.	Hydrochloric acid.	Sulphuric acid.	Nitric acid.
1.	Sodium [Potassium behaves in the same way].	Sodium + hydrochloric acid = sodium chloride + hydrogen.	Sodium + sulphuric acid = Sodium sulphate + hydrogen	Sodium + nitric acid = sodium nitrate + nitrogen peroxide + water.
2.	Copper ...	No action with dilute acid. In the presence of air, it slowly dissolves in the concentrated acid, forming copper chloride.	No action in the cold, but the hot and concentrated acid gives sulphur dioxide—Copper + sulphuric acid = copper sulphate + sulphur dioxide + water.	Acts readily even in the *cold* – Copper + *nitric acid* = copper nitrate + nitrogen peroxide + water
3.	Silver ...	No action ...	No action in the cold, but hot and *concentrated acid* gives sulphur dioxide Silver + sulphuric acid = silver sulphate + sulphur dioxide + water.	Acts readily even in the cold—Silver + nitric acid = silver nitrate + nitrogen peroxide + water.
4.	Zinc ...	Zinc + hydrochloric acid = zinc chloride + hydrogen. Practically no action on pure zinc.	Zinc + sulphuric acid (dil.) = zinc sulphate + hydrogen. Zinc + sulphuric acid (conc.) = zinc sulphate + sulphur dioxide + water.	Zinc + nitric acid = zinc nitrate + oxide of nitrogen + water.
5.	Mercury ...	No action ...	No action in the cold, but hot and concentrated acid gives sulphur dioxide—Mercury + sulphuric acid = mercury sulphate + sulphur dioxide + water.	Mercury + nitric acid = mercury nitrate + nitrogen peroxide + water.

6.	Aluminium	Aluminium + hydrochloric acid = aluminium chloride + hydrogen. It dissolves slowly in cold and dilute hydrochloric acid, but rapidly in the hot and strong acid.	Aluminium + sulphuric acid (dil). = aluminium sulphate + hydrogen. Aluminium + sulphuric acid (hot and *concentrated*) = aluminium sulphate + sulphur *dioxide* + *water*.	Nitric acid is without ny action on aluminium.
7.	Tin ...	Dilute acid has a slow action. Tin + hydrochloric acid (hot and conc.) = tin chloride + hydrogen.	Dilute acid has no action. Hot and conc. acid *gives* sulphur dioxide— Tin + sulphuric *acid* = tin sulphate + sulphur dioxide + water.	Tin + nitric acid (conc.) = metastannic acid + nitrogen peroxide + water. Absolutely pure nitric acid has no *action on* tin. Dilute nitric acid *slowly* dissolves tin, forming *tin* nitrate, ammonium nitrate and water.
8.	Lead ...	No action ...	Dilute acid has no action. Hot and conc. acid gives sulphur dioxide— Lead + sulphuric acid = lead sulphate + sulphur dioxide + water.	Lead + *nitric acid* = lead nitrate + *nitrogen* peroxide + water.
9.	Iron ...	Iron + hydrochloric acid = iron chloride + hydrogen.	Iron + sulphuric *acid* (*dil.*) = iron sulphate + *hydrogen*. Iron + sulphuric acid (*conc.*) = *iron* sulphate + sulphur dioxide + water.	Iron + nitric acid (dilute *and* cold) = iron nitrate + *nitrogen* peroxide + water.

CHAPTER XXV.

SYMBOLS, FORMULÆ AND EQUATIONS.

66. **Atoms and molecules.**—Matter, whether in the solid, liquid, or gaseous state is regarded by the scientists to be made up of extremely minute and distinct particles. A piece of stone can be ground into dust. A spoonful of water may be divided into a large number of drops. A substance may theoretically be divided and sub-divided into an indefinitely large number of fine particles. But in practice, a limit is reached beyond which a further sub-division is not possible. Scientists suppose matter to be made up of very small particles. These extremely minute particles, which are incapable of further division, are called **atoms**. An atom is the smallest particle of an element. It is extremely small in size, a million times smaller than the breadth of a hair, and cannot be seen even with the most powerful microscope.

It has now been found that atoms are not indivisible. They consist of still smaller particles known as **electrons.** But an atom cannot be split up into electrons by any known chemical process. Nothing less than an atom ever takes part in a chemical reaction, and it may therefore be regarded as a unit of chemical change.

Atoms cannot generally exist in the free state, and when liberated as such they unite to form 'compound atoms' or *molecules*. A **molecule** is the smallest particle of an element or a compound

which exists free in the gaseous state. It is the molecules that form the mass of matter. A molecule, so to say, is a little building of which atoms are the bricks.

Since molecules are collections of atoms, they can be divided into the latter. The molecules of an element are composed of atoms of the same kind, while those of a compound are made up of atoms of unlike kinds, because the smallest particle of a compound must consist of at least two different kinds of atoms. Thus a molecule of oxygen gas is found to contain two atoms of oxygen, whereas a molecule of hydrochloric acid gas is composed of one atom of hydrogen and one of chlorine.

The actual size of a molecule is very small indeed. To have some idea of the size of a molecule, imagine a single drop of water to be magnified to the size of the earth, each molecule in the drop being proportionately enlarged. Then each molecule will be about the size of a walnut.

67. **Symbols.**—Elements were formerly represented by fanciful figures. Thus gold was represented by the symbol ✲ , a symbol for the sun, silver by) , a symbol for the moon, and so on. But this system was found inconvenient and soon given up. The system now used is a kind of short-hand for writing the names of elements. The first letter of the English or Latin name of the element represents that element. For instance, H stands for hydrogen, O for oxygen and C for carbon. In cases where more than one element begin with the same letter, a second distinguishing letter is added. Thus carbon, calcium, chlorine, and copper (Latin:

cuprum) are represented by the symbols C, Ca, Cl and Cu respectively.

A symbol denotes not only the name of the element but also a definite quantity of the element which it represents. For example, O is not only an abbreviation for oxygen, but it also means 16 parts by weight of oxygen. Similarly, Cl signifies not only the element chlorine but also 35·5 parts by weight of chlorine.

A list of some important elements with the corresponding symbols is given below:—

Element.	Symbol.	Element.	Symbol.	Element.	Symbol.
Aluminium...	Al	Gold ... (*aurum*)	Au	Phosphorus ...	P
Antimony ... (*stibium*)	Sb	Hydrogen ...	H	Platinum ...	Pt
Arsenic ...	As	Iodine ...	I	Potassium ... (*kalium*)	K
Barium ...	Ba	Iron ... (*ferrum*)	Fe	Silver ... (*argentum*)	Ag
Bromine ...	Br	Lead ... (*plumbum*)	Pb	Sodium ... (*natrium*)	Na
Calcium ...	Ca	Magnesium ...	Mg	Sulphur ...	S
Carbon ...	C	Mercury ... (*hydrargyrum*)	Hg	Tin ... (*stannum*)	Sn
Chlorine ...	Cl	Nitrogen ...	N	Zinc ...	Zn
Copper ... (*cuprum*)	Cu				

68. Chemical formulæ.— When symbols of two or more elements are placed side by side, the resulting symbol means that the elements are in chemical combination. This combination of symbols is known as the *formula* and represents the composition of the compound for which it stands. A

formula is therefore a brief and convenient method of representing the composition of a compound as determined by experimental results. The compound formed by the combination of hydrogen (H) and chlorine (Cl) is represented by the formula HCl. Since the symbol H stands for 1 part by weight of hydrogen and Cl for 35·5 parts by weight of chlorine, the formula HCl means that 1 part of hydrogen combines with 35·5 parts of chlorine to form 36·5 parts of hydrochloric acid.

When more than one atom of an element is present in a compound, a numerical figure is attached to the right hand corner of the symbol to indicate the number of atoms present. Thus the formula of calcium carbonate is written as $CaCO_3$, which means that the compound contains 1 atom of calcium, 1 of carbon, and 3 of oxygen. It also signifies that 40 parts of calcium combine with 12 parts of carbon and $3 \times 16 = 48$ parts of oxygen to give 100 parts of calcium carbonate.

When a number is placed in the beginning of the formula of a compound, it represents the number of molecules of the compound, *e.g.* $2CaCO_3$ means two molecules of calcium carbonate. If, however, the formula contains more than one group of elements, the number multiplies only that group in front of which it is placed. For instance, the formula $CuSO_4, 5H_2O$ means that the compound contains 1 molecule of anhydrous copper sulphate ($CuSO_4$) and 5 molecules of water (H_2O). Sometimes the formula is put within brackets and the number placed outside the brackets. Ferric sulphate, for example, is represented by the formula $Fe_2(SO_4)_3$, which indicates that the substance

contains 2 atoms of iron, 3 of sulphur, and $4 \times 3 = 12$ of oxygen.

Thus we find that the formula of a compound gives: (*i*) the names of the elements present in a molecule of the compound, (*ii*) the number of atoms of these elements, and (*iii*) the relative weights of these atoms as represented by their symbols.

69. **Chemical equations.**—When a chemical action takes place and new substances are formed, the reacting substances and the products of reaction may all be represented by chemical formulæ. These formulæ are therefore very helpful in representing a chemical change. When zinc and sulphur combine together, they form zinc sulphide. This reaction may be represented as—

$$\underset{\text{zinc}}{Zn} + \underset{\text{sulphur}}{S} = \underset{\text{zinc sulphide}}{ZnS}$$

Similarly, the decomposition of potassium chlorate into potassium chloride and oxygen by heat may be expressed as—

$$\underset{\text{potassium chlorate}}{2KClO_3} = \underset{\text{potassium chloride}}{2KCl} + \underset{\text{oxygen}}{3O_2}$$

Again, when silver nitrate and sodium chloride react together, they form silver chloride and sodium nitrate, and the reaction may be represented as—

$$\underset{\text{silver nitrate}}{AgNO_3} + \underset{\text{sodium chloride}}{NaCl} = \underset{\text{silver chloride}}{AgCl} + \underset{\text{sodium nitrate}}{NaNO_3}$$

These brief expressions of chemical changes by means of formulæ are termed **equations.**

Each chemical equation must **correctly** express: (*i*) the nature and proportion of atoms and molecules taking part in a chemical action, (*ii*) pro-

portion by weight of the reacting substances, and (*iii*) relative volumes of gases concerned in the action. Since by the law of conservation of matter, matter can neither be created nor destroyed, the weight of the new products formed must be equal to that of the reacting substances. In other words, the number of atoms of each element, as represented by the symbols, must be equal on the two sides of an equation. For instance, sulphuric acid acts on zinc to give zinc sulphate and hydrogen, and the reaction is represented as—

$$\underset{\text{zinc}}{Zn} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{zinc sulphate}}{ZnSO_4} + \underset{\text{hydrogen}}{H_2}$$

The equation shows that 1 atom of zinc reacts with 1 molecule of sulphuric acid to form 1 molecule of zinc sulphate and 1 of hydrogen. The number of atoms of zinc hydrogen, sulphur and oxygen on the two sides of the equation are equal.

The formation of water from hydrogen and oxygen is expressed by the equation—

$$\underset{\text{hydrogen}}{2H_2} + \underset{\text{oxygen}}{O_2} = \underset{\text{water}}{2H_2O}$$

In addition to the nature and amount of reacting substances and the product formed, this equation tells us that 2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of water vapour.

A few more equations representing some important chemical reactions that we have studied in the book are:—

1. Preparation of carbon dioxide—

$$\underset{\text{calcium carbonate}}{CaCO_3} + \underset{\text{hydrochloric acid}}{2HCl} = \underset{\text{calcium chloride}}{CaCl_2} + \underset{\text{water}}{H_2O} + \underset{\text{carbon dioxide}}{CO_2}$$

2. Preparation of carbon monoxide—

$$CO_2 + C = 2CO$$

carbon dioxide	carbon	carbon monoxide

3. Preparation of chlorine—

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2$$

manganese dioxide	hydrochloric acid	manganese chloride	water	chlorine

4. Preparation of hydrochloric acid—

$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl$$

sodium chloride	sulphuric acid	sodium sulphate	hydrochloric acid

5. Preparation of sulphur dioxide—

$$Cu + 2H_2SO_4 = CuSO_4 + 2H_2O + SO_2$$

copper	sulphuric acid	copper sulphate	water	sulphur dioxide

6. Preparation of ammonia—

$$2NH_4Cl + Ca(OH)_2 = CaCl_2 + 2NH_3 + 2H_2O$$

ammonium chloride	calcium hydroxide	calcium chloride	ammonia	water

7. Preparation of nitric acid—

$$2NaNO_3 + H_2SO_4 = Na_2SO_4 + 2HNO_3$$

sodium nitrate	sulphuric acid	sodium sulphate	nitric acid

PANJAB UNIVERSITY QUESTION PAPERS.

Chemistry—Metals.

1927

1. What are the principal characteristics of metals? Do you consider the division of elements into metals and non-metals satisfactory?

2. Enumerate the chief properties and uses of any two of the following metals:—
   (*a*) aluminium, (*b*) tin, (*c*) silver, and (*d*) zinc.

1928

To what uses are the following metals put:—
(*a*) Mercury, (*b*) gold, (*c*) zinc, and (*d*) copper?
On what properties of the metals do these uses depend?

1929

1. Write a short account of the properties and uses of the following:—
   (1) Potassium permanganate, (2) saltpetre, (3) nitric acid, and (4) potassium bicarbonate.

2. To what uses are the following metals put:—
   (*a*) Aluminium, (*b*) iron, (*c*) zinc, and (*d*) silver?
   On what properties do these uses depend?

1930

Enumerate the chief properties and uses of any three of the following metals:—
Copper, lead, zinc and mercury.

Or,

Give the composition, properties and uses of (*a*) caustic potash, (*b*) sulphuric acid, (*c*) bronze, and (*d*) green vitriol.

1931

Mention the distinguishing characteristics and the important uses of the different varieties of iron. What is the effect of long exposure of iron to moist air? What devices are used to protect iron from the action of moist air?

Or,

Write what you know about the action of—
(*a*). Sulphuric acid on lead.
(*b*). Nitric acid on silver.
(*c*). Hydrochloric acid on copper.
(*d*). Water on zinc.

State clearly the conditions under which the above reactions would take place and the products of reactions obtained in every case.

1932

1. Compare the properties of copper with that of silver. What happens when hot and strong sulphuric acid acts on these metals ?

2. What happens when—
   (*a*). Ammonium nitrite is heated.
   (*b*). Hot and concentrated sulphuric acid is dropped on common salt.
   (*c*). Zinc dust is heated with caustic soda solution.

1933

1. Describe the manufacture of sodium carbonate.

Or,

Describe the occurrence, properties and uses of silver. What happens when sulphuric acid, nitric acid and hydrochloric acid act on the metal.

1934

1. Describe the occurrence, properties and uses of copper. Describe the action of common acids on it in the cold and in the hot.

2. Three bottles contain concentrated hydrochloric, sulphuric and nitric acids. How would you identify them without diluting the acids ?

1935

1. Give the chemical names and composition of the following :—
   (*a*) Blue vitriol, (*b*) green vitriol, (*c*) Epsom salt, (*d*) gypsum, and (*e*) alum.

Or,

What are alloys and amalgams ? Give examples of some common alloys and their composition.

1936

1. Give the composition, properties and uses of—
   (*a*) Washing soda, (*b*) caustic soda, (*c*) potassium chlorate, and (*d*) brass ?

Or,

To what uses are the following metals put :—
(*a*) Copper, (*b*) silver, (*c*) iron, and (*d*) aluminium ?
On what properties of the metals do these uses depend ?

APPENDIX

SCIENTIFIC HOBBIES.

A number of Scientific hobbies could be started in schools or colleges to arouse the interest of the students in vocational training. It is not possible for any one school to take up more than a few hobbies. Every school or college will have to make its own selection, depending upon the local conditions and facilities provided.

Besides effecting some saving in our daily expenses, these hobbies would keep our young men away from idle habits and bad company. They would also stimulate an active interest and enthusiasm for small-scale industries which are so badly needed in the present days to solve the problem of unemployment. It will be advisable to concentrate on only those hobbies which require no elaborate apparatus, incur small initial and recurring expense, and can bring a fairly decent living to an educated young man.

It is not possible to give here more than a brief outline of some of the very simple and cheap hobbies that would be found useful as well as interesting. For further and detailed information the teachers and the students are requested either to refer to the issues of the Punjab Educational Journal or to write to the authors of the book.

I. Soap-making.

Soap is an article of daily use. Good washing soap could be easily prepared at home or in the

school by the following process :—

Cocoa-nut oil (*naryal ka tel*)	...	$3\frac{1}{2}$ parts
Mohwa oil (*mowe ka tel*)	...	$3\frac{1}{2}$ parts
Caustic soda flakes	...	1 part
Water	...	4 parts

Mix the two oils together and heat gently, if necessary, to melt them. In the meanwhile, dissolve the soda in water and let the solution, *soda lye*, cool. When the mixture of oils and the lye are at about the same temperature, which can be tested by putting a finger in each of them, add the lye to the oil mixture, stirring the mass well with a wooden rod.

When the whole lye has been added and a uniform pasty mass obtained, transfer it into an iron mould. Cover the mould with a wooden board or a gunny bag and leave it overnight. Next morning cut it into cakes of the desired size.

This gives a pure and good washing soap which, when slightly coloured and scented, could be used for ordinary toilet purposes.

Precautions.—Caustic soda is very corrosive and should not therefore be touched with the hand. Its solution should be prepared in an iron pan (*karahi*) by stirring with an iron ladle (*khaunchi*).

For full information regarding various kinds of soap the student is advised to read any of the two books :—

1. Soap making by the Cold Process, published by the Government Printing Works, Madras.

2. Manufacture of Soap, published by the Industry Book Depot., Calcutta.

II. Ink-making.

The common writing ink powders or tablets are a mixture of dye-stuff (*e.g.* methylene blue,

eosine, naphthol green, or nigrosine) with about half the weight of gum arabic (*kikar ki gond*). A good writing ink may be prepared as follows:—

1. *Blue-black Ink.*

Gall-nuts (*maju*)	... 50	parts
Iron sulphate (*hira kasis*)	15	,,
Gum arabic	5	,,
Aniline blue soluble	5	,,
Water	... 800	,,
Carbolic acid	... 1	,,

Crush the gall-nuts, soak them in water, and let them ferment for about 10 days. Add water to replace the one last by evaporation every day. Boil the liquid and strain through a muslin cloth. Repeat the process twice with the residue left on the cloth.

Cover the powdered gum arabic with 100 parts of water and let it stand overnight. Dissolve iron sulphate in about 150 parts of water and add it to the gall extract. Now add the gum solution and aniline blue. Stir well and make the water to 800 parts. Finally add carbolic acid and transfer it into a stoneware or glass bottle. Leave it for a few days and pass it through a filter sheet.

2. *Red Ink*

Eosine	...	20 parts
Gum arabic	...	10 parts
Water	...	1,000 parts
Carbolic acid	...	2 parts

Add the gum to 100 parts of water and let it stand overnight. Dissolve eosine in 500 parts of hot water. When it has dissolved completely, add the gum solution. Stir well and add the remaining quantity of water. Finally add the carbolic acid.

For detailed information read 'Ink Manufacture by Sigmund Lehner.'

III. Chalk-making.

The main ingredient of black-board chalk is plaster of Paris obtained by burning gypsum in a special kiln. The sticks are obtained by pouring the pasty mass in a brass mould consisting of six grooved brass plates. Each plate has got about 24 grooves, and when the plates are joined together, the two opposite grooves unite to form a hole of the size and shape of an ordinary chalk stick. Each mould costs about Rs. 40/- to 50/- and could be prepared by any of the firms dealing in black-board chalk at Gujranwala.

Mix plaster of Paris with 2 to 3 per cent. of white clay. Add a trace of ultramarine blue (*dhobi ka nil*) to water, and pour this water to the above mixture so as to form a thin pasty mass. Transfer the mass gently into the mould and leave for a few minutes. Separate out the plates carefully. Chalk sticks, with smooth surface, are obtained. Let them dry in the sun, and then collect in a box.

Precautions.—1. Use *fresh* plaster of Paris which, if exposed for some time to air, loses the property of setting.

2. Clean the plates and oil them with a mixture of 3 parts of kerosene oil and 1 part of rape-seed oil (*karwa tel*) before pouring the pasty mass into the holes of the mould.

IV. Refining of Hair Oils.

Most of the oils in the market are prepared from white oil (*i. e.* refined kerosene oil) which is injurious both to the skin and the hair. Rape-seed

oil (*sarson ka tel* or *karwa tel*) is generally disliked due to its dirty colour and unpleasant smell. Fairly good sesame oil (*tili ka tel*) is available in almost every town. If treated with about 10 per cent. weight of fuller's earth (*pili matti*), chalk, kieselguhr, or animal charcoal and then filtered through the same substance, a yellowish coloured oil is obtained. A little of any scent, *anwala* or *moulsari*, when added to this oil, will give a pure vegetable oil intended for daily use.

Cocoa-nut oil, when subjected to a similar treatment, gives a good hair oil. But as it solidifies in winter, it can only be used in summer. If necessary, the oil may be coloured red with alkanet root (*ratan jot*) or green with *neem* leaves.

The oil thus prepared does not look so decent as those placed in beautiful bottles in the market. But it is certainly very much cheaper and better than the latter, because it is a pure vegetable oil.

The refining of oils with sulphuric acid or caustic soda requires some technical knowledge and skill which are beyond the scope of the present discussion.

V. Boot Polishes.

A shoe cream differs from a polish only in the quantity of turpentine oil. The latter contains less of the oil and is therefore comparatively a stiff mass.

Brown Boot Polish

Montan wax (refined)	...	3 parts
Ceresine wax	...	1 part
Stearin	...	1 ,,
Carnauba wax	...	1 ,,
Potassium carbonate	...	0·5 ,,
Phosfine brown	...	0·5 ,,

Dissolve the potassium carbonate in about 5 parts of water. Melt the waxes together by gentle heating and add the carbonate solution with constant stirring. Now add phosfine brown dissolved in a little water till a thick pasty mass is obtained. Stop heating, let the mass cool, and dilute with turpentine oil to the required consistency.

For dark-brown colour, Bismark brown may be used in place of phosfine brown.

Black Boot Polish

Montan wax (raw)	...	3 parts
Ceresine wax	...	1 part
Stearin	...	1 ,,
Carnauba wax	...	1 ,,
Paraffin wax	...	1 ,,
Potassium carbonate	...	0·5 ,,
Nigrosine	...	0·7 ,,

Proceed as in the case of brown boot polish.

Note. For high brilliancy take more of harder waxes like montan, carnauba, or paraffin wax, and dilute with more of turpentine oil.

VI. Tooth Powder.

A mixture of finely powdered charcoal and common salt (or alum) is said to form a good tooth powder. But the following recipe will give a better and more useful dental powder:—

Precipitated chalk	...	13 parts	
Magnesium carbonate	...	7 ,,	
China clay (fine)	...	1 part	
Saccharine	...	Trace	
Thymol (*sat ajwain*)	...	0·2 ,,	4 per cent. of the above ingredients
Camphor (*kafur*	...	0·15 ,,	
Peppermint oil	...	0·5 ,,	
Cinnamon oil (*darchini ka tel*)		0·15 ,,	

Mix the first four ingredients in a mortar, then add thymol and camphor. Powder and pass through a muslin cloth. Finally add the mixture of the other ingredients.

For carbolic tooth powder, omit the last three ingredients and add carbolic acid, 2 per cent. of the weight of the whole mixture.

A little 'pigment pink' may also be added to impart an attractive appearance to the powder.

VII. Polish for Furniture.

Good furniture, if polished every now and again, looks decent and lasts fairly long. Polish for furniture may be prepared in the following manner:—

Methylated spirit	...	1 bottle (24 oz.)
Shellac granular (*lakh dana*)	...	1½ chhk.
Mustic (*mastagi*)	...	1 tola
Sandras (*sundras*)	...	½ tola

Powder the last three ingredients and put them in the methylated spirit contained in a bottle. Cork the bottle and leave it overnight, or better keep it in the sun for a few hours. Strain through cloth and apply with a pad, *i.e.* a muslin piece with cotton wool.

For *shisham* wood polish, add ½ tola of dragon's blood (*khun-i-shoshan*).

In order to prepare the surface for polish, rub it well with the sand-paper.

VIII. Silvering Glass.

The process of silvering glass consists of three separate steps:—

I. *Cleaning the surface.*—The glass plate to be silvered is thoroughly cleaned by washing first with sodium carbonate solution (or on a commercial scale

with finely powdered bath-brick) in order to remove the oily or greasy matter, if any present, then with tap-water, and finally with distilled water.

II. *Silvering the surface.*—The cleaned plate is then levelled up on a wooden frame, or preferably on a warm table, and washed with a solution of stannous chloride in order to obtain a uniform and coherent deposit of silver.

The silver solution, which is prepared as given below, is now poured on the prepared plate and allowed to stand for about half an hour. In practice, on a large scale, the temperature is kept at 40°C., when a good film of silver is deposited in about 20 minutes.

III. *Fixing silver on the plate.*—In order to protect the silvered surface from the action of sulphuretted hydrogen, which is always present in minute traces in the atmosphere, it is coated with shellac varnish and then painted with red lead (*sandhur*) or red oxide of iron (*geru*).

Preparation of the Silver Solution.—This consists of two separate solutions A and B, kept in stock and used when required.

Solution A.—Dissolve 10 grams of silver nitrate in 80 c.c. of distilled water. Add ammonia solution to it drop by drop till the precipitate formed just dissolves. Dilute to a litre by adding distilled water, and finally filter.

Solution B.—Dissolve 2 grams of silver nitrate in as little distilled water as pssible. Add this to a litre of boiling water. While boiling, add 1·66 grams of Rochelle salt (sodium potassium tartrate) and continue boiling for 20 minutes or more until the liquid is very clear, leaving a grey precipitate behind. Filter while hot, and make up to one litre by adding distilled water.

When required for use, mix the two solutions and pour the mixture quickly over the prepared surface, as described above.

Precautions.—1. Keep the two solutions in dark-coloured bottles away from strong or direct sunlight.

2. Carry out the process of silvering in a room where there is no direct sun.

IX. Washing and Stain Removing.

1. WASHING.

Our washermen often spoil our expensive clothes and suits due to their ignorance or carelessness. If silk and woollen clothes could be washed at home, it would add considerably to their life and decency.

For silk garments, take half a tea cupful of soap jelly (prepared by dissolving 1 part of Sunlight, Jallo, or any other neutral soap in 3 parts of boiling water) in an enamelled basin. Add water (about 6 times the weight of the cloth) to it, soak the cloth in it, and rub well. Now wash the cloth in cold water.

Take 1 dessert spoonful of citric acid (*khatte ka tezab*) and 1 tea spoonful of gum solution, and dissolve them in lukewarm water which is about 6 times the weight of the cloth. Work the cloth for about 5 minutes in this solution. Wash with ordinary water, squeeze and dry. If lustre is required, iron the garment in a semi-dry condition.

For woollen clothes, take a cupful of soap jelly, 2 tea spoonfuls of ammonia liquor and water (about 8 to 10 times the weight of the cloth) in a washing basin, wet the cloth with water, introduce in this soap solution and work well for about 10 minutes. If some spots are visible, pour some more soap and jelly over

them, and rub gently. Wash with ordinary water. Now put 1 tea spoonful acetic acid in water (about 8 times the weight of the cloth) and work the material again for a few minutes in this solution. Wash, squeeze and dry in the shade, turning the garment inside out. While ironing the cloth, cover it with a moist piece of muslin.

In case of white woollen garments, *e.g.* flannel pants, use citric acid (about 1 oz. per pant.) in place of acetic acid.

The following recipe of a **cleansing fluid** will be found very useful in washing most of the silk and woollen garments:—

Take Sunlight or any other neutral soap and cut it into thin shavings. Boil the weighed quantity of soap in water in an enamelled dish till it has completely dissolved. Stop heating.

Soap	...	6 parts
Ammonia	...	4 ,,
Glycerine	...	2 ,,
Methylated spirit		4 ,,
Water	...	100 ,,

Let the liquid cool, and add the remaining ingredients and keep in a stoppered or corked bottle. This fluid can also be used for washing a garment after the stains have been removed from it.

II. REMOVAL OF STAINS.

Stains are of many kinds, and it is sometimes impossible to remove them without injury to the garment. All stains are more easily removed when fresh.

While removing a stain, hold the stained portion of the cloth in a hoop or stretch it across a support

of some kind. Place a pad of blotting-paper or muslin immediately under the stain. Avoid water unless specially required. Use a small quantity of the reagent with a medicine dropper. Just as the stain disappears, dry the cloth with a fresh blotting-paper or with a hot iron.

Treat the spot on both sides of the cloth, otherwise the removal may be superficial, and the stain may reappear.

Some Common Stains.

1. *Ink* (blue-black) (*i*). *From Cotton or Wool.*—Treat the spot with 4 per cent. oxalic acid solution for about 5 minutes. Squeeze, dip and rub in 5 per cent. solution of sodium hydrosulphite. If the spot is not removed, use hot solutions and repeat the process.

(*ii*). *From Silk.*—Dip the stained portion in 4 per cent. oxalic acid solution as in (*i*). Boil in 5 per cent. Rongalite C (*i.e.* sodium sulpho-oxalate) solution. Repeat the process till the stain is removed. Finally wash with cold water.

2. *Oil.*—For cotton, use 10 per cent. of solution of soda-ash in hot water, and for silk and wool treat with cold and dilute solution of ammonia.

3. *Grease.*—Pour some oil of turpentine or carbon tetrachloride on the spot and rub well. If the stain persists, use the hot liquid, and then wash with soap and water.

For cycle or medicine grease, treat the stain with petrol. Add fresh petrol and rub the stained portion well with it. Repeat and finally wash with soap and ammonia.

4. *Turmeric.*—This can be easily removed from cotton on treatment with hot solution of soda-

ash or soap. This is the usual procedure adopted by our washermen. But in case of silk or wool, the following treatment is more effective.

If fatty matter is also present, as is often the case, remove the fat with dilute ammonia. The colour of the stain changes to reddish-brown, because alkalies turn turmeric brown in colour. Wash with water and place the stained portion in 1 per cent. potassium permanganate solution. Squeeze and now put it in 8 per cent. sodium bisulphite solution. Finally wash thoroughly with water.

5. *Blood.*—Treat the spot at first with 4 per cent. oxalic acid solution, then with 1 per cent. potassium permanganate, and finally with 8 per cent. sodium bisulphite solution.

In case of silk, blood stains are difficult to remove. Spot with 50 per cent. lactic acid solution, then dip in hydrogen peroxide. For old stains, apply a paste made up of the yellow and white of an egg and a few drops of dilute sulphuric acid, before the above treatment.

Proceed in the same way in case of beetle (*pan*) stains. In certain cases, rubbing with an onion or application of mustard water removes these stains.

6. *Milk, Coffee, or Tea.*—Use 5 per cent. solution of borax at about 60°C. In case of old stains, apply a mixture of the white of an egg and glycerine for about ten minutes, and then treat with the borax solution.

7. *Paint and Varnish.*—Dip the stained portion in hot oil of turpentine, and rub well. Repeat the process, if necessary. If the stain persists, rub in a little grease or aniline oil. Now wash with soap and ammonia. Remove the final traces with 5 per cent. solution of sodium hydrosulphite.

8. *Boot Polish.*—Treat the spot with hot oil of turpentine. Repeat if necessary. Wash with soap solution. If some traces are left, use 5 per cent. sodium hydrosulphite solution.

9. *Iron Rust.*—Dip in a 5 per cent. solution of oxalic acid at about 60°C and rub well. Now treat with 1 per cent. solution of bleaching powder and acetic acid in case of cotton, or 5 per cent. sodium hydrosulphite solution in case of silk or wool. Finally wash with water.

10. *Fruit Juice.*—Wash with water to remove sugary matter and then treat with ammonia to neutralise the acids present in the juice. Now remove the colouring matter with the solution ot bleaching powder or sodium hydrosulphite as given in 9.

11. *Iodine.*—Use dilute ammonia solution for silk and wool and 50 per cent. soda-ash solution for cotton. If the spot is not removed, use 5 per cent. hypo (sodium thiosulphate) solution.

12. *Silver Nitrate.*—In case of caustic lotion, protargol, argyrol, or any other silver stains, treat with 5 per cent. potassium cyanide solution. Finally wash with soap and water.

Gold thread (*silma stara*) often becomes dark-brown or black due to oxidation. The original brightness can be restored by rubbing the thread carefully with a crystal of potassium cyanide held in a muslin piece and moistened with water.

Note : Potassium cyanide is a deadly poison and it must be handled with very great care. No student should be allowed to use its solution.

Spotter's cabinet.—A small box with 12 reagent bottles, popularly known as the 'spotter's cabinet',

could be prepared within a few rupees and used for removing most of the common stains from our garments. The box contains the following reagents:—

1. Oxalic acid, 4 per cent. solution.
2. Sodium hydrosulphite, 5 per cent. solution.
3. Soda-ash, 10 per cent. solution.
4. Dilute ammonia.
5. Potassium permanganate, 1 per cent. solution.
6. Sodium bisulphite, 8 per cent. solution.
7. Petrol.
8. Turpentine oil.
9. Borax, 5 per cent. solution.
10. Bleaching powder, 1 per cent. solution.
11. Potassium cyanide, 5 per cent. solution.
12. Cleansing fluid.

The above chemicals could be had from Messrs. Beli Ram Bros., Lahore, or Havero Trading Co., Bombay.

X. Family Requisites.

1. Zinc lotion

Zinc sulphate	...	2 parts
Boric acid	...	10 ,,
Water (distilled or rose)		440 ,,.

Dissolve the two powders separately in hot water. Mix the two solutions and make up the required volume of water.

2. *Zinc Ointment*

Boric acid	...	33 parts
Zinc oxide	...	80 ,,
Vaseline	...	440 ,,

Mix the first two ingredients thoroughly with vaseline in a mortar or with a spatula on a porcelain tile. This gives a good ointment for small ulcers.

3. *Tincture Iodine*

Iodine	...	24 parts
Potassium iodide	...	24 ,,
Water	...	22 ,,
Methylated spirit	...	88 ,,

Dissolve potassium iodide in water and add iodine to the solution. Now add the methylated spirit.

For tincture iodine required for internal use, rectified spirit is used instead of methylated spirit.

4. *Lotion for Burns*

For small burns *carron oil*, prepared by mixing lime-water with an equal volume of any vegetable oil, *e. g.* linseed or cotton-seed oil, is generally employed. If, however, the burn covers rather a large portion of the skin, the following treatment will be found more effective—

Tannic acid	...	10 parts
Water	...	440 ,,

Dissolve the acid in water. Apply the solution freely to the burn, and then apply the following ointment—

Boric acid	..	3 parts
Vaseline	...	44 ,,

Mix the acid thoroughly with vaseline before use.

5. *Treatment for Cuts*

If the cut is very slight, wash with methylated spirit and apply tincture iodine. If there is bleeding, apply a saturated solution of ferric chloride to coagulate the blood and then tincture iodine. Bandage with a pad of gauze over the wound.

In case foreign matter like dirt or glass pieces are present in the cut, remove them cerefully and then apply tincture iodine. Be careful that the iodine reaches all crevices of the wound.

For severe cuts, bandage tightly to stop the flow of blood, apply hydrogen peroxide or ferric chloride solution to coagulate the blood, and send for the doctor.

INDEX

[*Vernacular names are given in Italics*].

A

Abdari ... 194
Acid, definition of the term ... 44
„ hydrochloric ... 102, 197
„ nitric 134, 198
„ sulphuric ... 124, 197
Acids 68
Acidic oxides ... 44
Action of acids on metals ... 197
Air ... 29
„ action of phosphorus on ... 30
„ a mixture 32
„ composition of ... 33
„ other gases in ... 33
Alabaster 128
Alkalies 69
Allotropy 117
Allotropic modifications 118
Alloys 160
Aluminium 184
Amalgams 159
Ammonia 131
Ammonium salts ... 133
Amorphous carbon ... 76
Anglesite 189
Animal charcoal ... 78
Annealing (of steel)... 194
Anode 63
Aqua fortis 134
Aqua regia 138

B

Baking soda 168
Base, definition of ... 69
Bases 69
Basic oxides 44
Bell metal 160, 188
Black-lead 75
Bleaching action of chlorine 108
Bleaching action of sulphur dioxide ... 122
Bleaching powder ... 110
Blue vitriol 128, 176
Bone-ash 143
Bone-black 78
Brass (*pital*) 160, 175
Brimstone 133
Bronze (*kansi*) ... 160, 176
Bujha hua chuna ... 91
Bunsen burner ... 97

C

Calamine 180
Candle-flame 98
Carbides 80
Carbon, allotropic forms of 74
Carbon, amorphous... 76
——, oxides of ... 83
Carbonado 75
Carbon dioxide ... 83
Carbonic acid snow... 87
Carbon monoxide ... 91
Carbonates 89
Cast iron 194
Catalytic agent ... 39
Cathode 63
Caustic lotion ... 178
Caustic potash ... 169
Caustic soda 163
Cement 91
Charcoal, animal ... 78
——, wood 76
Chemical action ... 24

Chemical changes ... 3
Chemical compound 20
Chemistry, object of... 2
Chili saltpetre ... 134, 170
Chlorides 107, 110
Chlorine 106
Chlorine water .. 107
Clay iron stone ... 193
Coal 76
Coal-gas 80
Coke 80
Combustible substance 95
Combustion ... 95
——, and flame ... 95
——, supporter of ... 95
Common salt 141
Composition of water 62
Copper 174
——, alloys of ... 175
Copper glance ... 174
Copper pyrites ... 113, 174
Copper sulphate ... 176
Crystallisation ... 13
Crystals 13
Cuprite 174

D

Davy's safety lamp ... 100
Decantation 10
Dehydrating agent ... 149
Deliquescent... ... 54
Destructive distillation of wood 76
Diamond 74
Distillation 14
Drinking water ... 49
Druggist's still ... 16
Dutch metal ... 160

E

Efflorescence ... 167
Electrodes ... 63
Elements and compounds 19
——, list of ... 20
Epsom salts 113
Eudiometer tube ... 64
Evaporation 8

F

Filtration 11
Flame 96
——, luminosity of ... 96
——, structure of ... 98
Flash point 100
Flowers of sulphur . 115
Fur in kettles ... 51

G

Galena 113,189
Galvanised iron ... 180,195
German silver ... 160,176
Glauber's salt ... 128,168
Graphite 75
Green vitriol 128
Gun metal 160
Gypsum 113,128

H

Hardening (of steel)... 194
Hardness of water ... 49
Hæmatite 193
Heavy spar 113
Hira kasis 128,196
Hobbies Scientific ... Appdx. i
Horn silver 177
Hydrochloric acid ... 102
Hydrogen 56
——, properties of ... 59
Hygroscopic ... 54

I

Ignition temperature... 99
Indestructibility of matter ... 4
Indestructibility of matter, Law of ... 6

Iron ... 193
——, cast ... 194
——, pyrites ... 113, 193
——, sulphate (*hira kasis*) ... 196
——, wrought ... 194

K

Kajal 79
Kankar 90
Khari 128, 168
Koh-i noor 74

L

Lampblack 79
Law of conservation or indestructibility of matter 5
Lead 189
—— alloys of 191
Leblanc process ... 166
Lime 90
Limestone 89
Litharge (*murda sang*) 189
Loha 193
Luminosity of flame... 96

M

Magnalium 186
Magnetite 193
Manganese dioxide ... 39
Mansal 26
Matter 1
——, kinds of ... 1
——, constitution of ... 1
Mechanical mixture ... 19
Mercury 181
Metals 157
——, characteristic properties of ... 157
Metals, distinction from non-metals ... 159
Minerals 159
Mixtures and compounds, distinction 19
Mortar ... 91
Murda sang 189

N

Namak 165
Nascent state... ... 109
Natural waters ... 47
Neutralisation ... 71
Nila tutia 128
Nitrates 138
Nitre 135, 170
Nitric acid 134
Nitrogen 36
Noshadar 17, 131

O

Octahedral sulphur ... 116
Oil-gas 81
Ores of metals ... 159
Oxidation 43
Oxides 43
——, acidic 44
——, basic 44
——, neutral 44
Oxides of carbon ... 83
Oxygen 38
——, uses of 44

P

Permanent hardness 51
Pewter 162
Phitkari 13
Phosphates 150
Phosphites 149
Phosphoric acid ... 149
Phosphorus 143
——, red (or amorphous) 146
——, yellow 147
Phosphorus pentoxide 149
——, trioxide ... 143
Physical and chemical changes 2
Plaster of Paris ... 128
Plastic sulphur ... 117
Platinised asbestos ... 123

Plumbago 75
Point of ignition ... 99
Potassium 169
Potassium nitrate ... 138, 170
Primus stove 97
Prismatic sulphur ... 116

Q

Quicklime 91

R

Rain-water 47
Red lead 189
Reduction 62
Respiration 87
River-water 48
Roll sulphur 115
Rusting of iron ... 31, 195

S

Sajji 166
Saltpetre 135
Salts 71
Sandhur 191
Saturated solution ... 7
Sea-water 48
Scientific hobbies ... Appdx. i
Selenite 128
Shora 135, 170
Silver 177
Silver glance 177
——, nitrate 178
Slaked lime 91
Sodium 162
Sodium amalgam ... 57, 163
——, bicarbonate ... 168
——, carbonate ... 166
——, hydroxide ... 163
——, nitrate ... 168
——, sulphate ... 168
Softening of water ... 51
Solder 160, 188
Solubility 8
Solute 7
Solution 7
Solvay or ammonia-soda process ... 167
Solvent 3
Spirit of hartshorn ... 131
Spring-water 47
Steel 194
Sublimate 17
Sublimation 17
Sufeda 191
Sugar of lead 191
Sulphates 127
Sulphides 118
Sulphites 122
Sulphur 133
——, extraction of ... 114
——, flowers of ... 115
——, monoclinic ... 116
——, octahedral ... 116
——, plastic 117
——, prismatic ... 116
——, roll 115
——, rhombic 116
Sulphur dioxide ... 121
Sulphur trioxide ... 123
Sulphuric acid ... 124
Sulphurous acid ... 122
Supporter of combustion 95
Synthesis 56

T

Tamba 174
Tempering (of steel) ... 195
Temporary hardness 50
Tin 187
Type metal 160, 191

V

Verdigris 175
Voltameter 63

W

Washing soda 166
Water 46
——, composition ... 62
——, drinking 49
——, filtration of ... 11
——, hard and soft ... 49
——, of crystallisation 53, 176
——, natural 47
——, rain 47

Water, river ... 48
——, sea 48
——, soft 50
——, spring 47
Waters, mineral ... 48
Welding 194
White lead 189, 191
——, vitriol 181
Wood charcoal ... 76
Wrought iron ... 194

Z

Zangar 175
Zinc 180
——, alloys of 180
Zinc blende 131, 180
Zincite 180
Zinc lotion 181
Zinc oxide 181
Zinc sulphate ... 181

104